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— SECTION B —

CHEMICAL SCIENCES

Contents

Page

Photosensitized Reactions of Carbon Tetrafluoride—*J. R. Dacey and J. W. Hodgins* - - - - - - - - - - - - 173

Dielectric Dispersion Studies of Polyvinyl Acetate in Toluene—*B. L. Funt and S. G. Mason* - - - - - - - - - 182

The Diffusion of Polystyrene in Different Solvents—*P.Z. Adelstein and C. A. Winkler* - - - - - - - - - 194

Synthesis of Amino Acids from Substituted Cyanoacetic Esters —*Paul E. Gagnon, Jean L. Boivin, and Paul A. Boivin* - - 207

Comparison of the Nitrolysis of 1,3-Dicyclohexylimidazolidine with that of Hexamethylenetetramine—*J. L. Boivin and George F Wright* - - - - - - - - - - - - - 213

Acid-Base Reactions in Nonaqueous Media Applicable to Analytical and Radiochemical Separations. I. Reactions in Organic Derivatives of Ammonium Chloride—*K. Starke* - 225

Synthesis in the Series of *Erythrina* Alkaloids—*K. Wiesner, Frank H. Clarke, and S. Kairys* - - - - - - - - - 234

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PHOTOSENSITIZED REACTIONS OF CARBON TETRAFLUORIDE[1]

BY J. R. DACEY[2] AND J. W. HODGINS[3]

Abstract

Hg (6^3P_1) at room temperature and at 300° C. and Hg (6^1P_1) at room temperature fail to react with carbon tetrafluoride at a measurable rate. Xe ($^3P_1$) causes carbon tetrafluoride to decompose with a quantum efficiency of about unity, to yield fluorine and an unidentified solid product. It is concluded that the energy necessary to break the first C–F bond in CF_4 is more than 154 and less than 194 kcal. per mole. Hydrogen atoms produced from Hg (6^3P_1) at room temperature and at 300° C. and from Hg (6^1P_1) at room temperature do not react with carbon tetrafluoride. It is concluded from this that the activation energy of the reaction $CF_4 + H \rightarrow CF_3 + HF$ is not less than 14 kcal. per mole.

I. Introduction

Since the fluorocarbons resemble the hydrocarbons in many respects, it is of interest to investigate their decomposition by the methods of photosensitization which have proved so successful in studying the elementary reactions of the hydrocarbons. This paper deals with some experiments on the photolysis and photosensitized decomposition of the first member of the fluorocarbon series, carbon tetrafluoride.

There are several methods of estimating the average energy of the C–F bond in CF_4. For example, one may start with the value for the heat of combustion of Norit carbon in fluorine, found by Ruff and Bretschneider to be 183.5 kcal. per mole (9). Using this value, along with the value for the energy of the F–F bond and that for the heat of sublimation of graphite, the heat of formation of carbon tetrafluoride can be calculated using Hess's law. The values obtained are, to say the least, uncertain. The heat of combustion of graphite in fluorine may differ widely from that of Norit carbon. Also, the dissociation energy of fluorine has been reported to be as low as 55 kcal. by Gaydon (5), and as high as 70 kcal. by Bodenstein, Jockusch, and Shing Hou Chong (2). The value of the heat of sublimation of graphite, too, is still controversial, having been assigned values of 125 to 170 kcal. per mole (6, 10). Accordingly, depending upon which data one favors, the average C–F bond energy in CF_4 may be as low as 104.6 or as high as 123.4 kcal. per mole.

[1] *Manuscript received November 19, 1949.*
Contribution from the Defence Research Chemical Laboratories, Ottawa, Canada, and the Royal Military College of Canada, Kingston, Ont. The laboratory work was carried out at the Defence Research Chemical Laboratories, and is also reported as D.R.C.L. No. 37.
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In methane, while the average bond energy is given by Pauling (8) as 87 kcal. corresponding to the low value for the heat of sublimation of graphite, the strength of the first C–H bond has been shown by several methods to be about 102 kcal. (11). By analogy, then, the energy required to remove the first fluorine atom from carbon tetrafluoride is probably greater than 130 kcal. Accordingly, it was expected that the 2537 Å mercury resonance lamp might not produce decomposition and that recourse would be necessary to the higher energy mercury radiation at 1850 Å, corresponding to 154 kcal. per mole of quanta. One objection to a mercury photosensitized decomposition is that solid mercury fluoride deposits on the cell windows, making it difficult to estimate the light intensity. For this reason, the study of the xenon photosensitized decomposition of the fluorocarbons is attractive, since the formation of solid fluorides is unlikely. This paper describes experiments with carbon tetrafluoride, using mercury atoms activated by the 2537 Å and 1850 Å resonance radiation, and xenon atoms activated by the 1470 Å xenon resonance line.

II. Apparatus and Techniques

The 2537 Å Apparatus

The apparatus used for the experiments with Hg 6^3P_1 is illustrated in Fig. 1. It consisted of a fused silica reaction cell (*A*), 4.5 cm. in diameter and 10 cm. long, connected to a manifold through the mercury cutoff (*B*). The gas under investigation was circulated through the cell and the two liter volume (*C*) by means of the mercury diffusion pump, which also served to keep the gas in the reaction cell saturated with mercury vapor. The pressure was measured

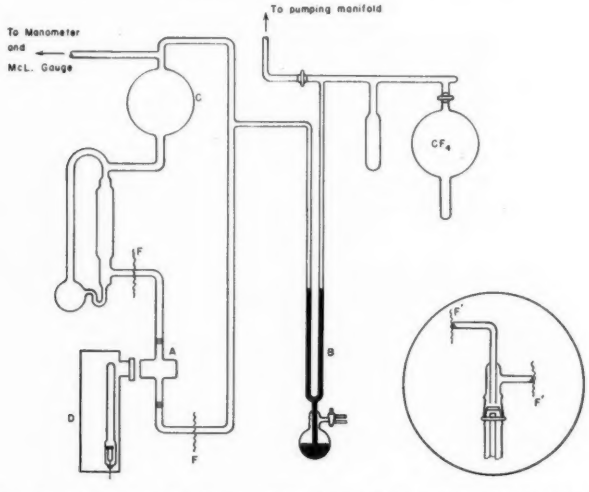

FIG. 1. *Apparatus for photosensitization with Hg* (6^3P_1) *and Hg* (6^1P_1).

by means of a manometer, using a cathetometer, or by means of a McLeod gauge. For the experiments at high temperatures the cell was enclosed in a heater which consisted of an insulated ceramic tube wound with resistance wire. The temperature was measured on the face of the cell by means of a thermocouple of fine wires.

The reaction cell was illuminated by a mercury resonance lamp housed in a transite box (D) with a window 2 in. in diameter. The lamp consisted of a 20 in. quartz tube ½ in. in diameter bent in the form of a U. The intensity of 2537 Å radiation entering the cell was determined by chemical actinometry using uranyl oxalate, and by means of a Kipp thermopile. The values obtained by both methods were in good agreement, the intensity being 2.8×10^{-8} einsteins per second.

In using the apparatus, the desired amount of reacting gases was admitted through the manifold, the cutoff closed, and the pressure accurately noted. The temperature of the two liter volume was then observed. The mercury pump and if necessary the heater were turned on. After thermal equilibrium had been reached the pressure was again recorded and the lamp turned on. The pressure was measured from time to time throughout the run. When the run was finished, the lamp, heater, and pump were turned off. After the whole apparatus had returned to room temperature the pressure was again measured and compared with the initial pressure reading.

The 1850 Å Apparatus

The apparatus for the experiments with Hg (6^1P_1) was identical with that described above except for the lamp and reaction cell, which are illustrated in the insert to Fig. 1. The lamp which produced 9.5×10^{-9} einsteins per second is described in detail elsewhere (4).

Experiments at room temperature only were made with the 1850 Å lamp. The procedure was the same as that described for the 2537 Å lamp.

The 1470 Å Apparatus

The apparatus for the experiments with Xe ($^3P_1$) is illustrated in Fig. 2. The resonance lamp which produced 1.9×10^{-9} einsteins of 1470 Å per second is described in detail elsewhere (4). The discharge tube is surrounded by the reaction space through which the gas is circulated by the pump P_2, a magnetically operated vibrating reed. The reaction space and lamp are immersed in a bath through which water circulated at room temperature. The Toepler pump is used to remove samples of the reaction mixture for analysis with the mass spectrograph. The pressure is measured by means of a McLeod manometer.

The reaction space is connected to a manifold by the stopcocks A and B. The connecting space between A and B, which includes the manometer, has

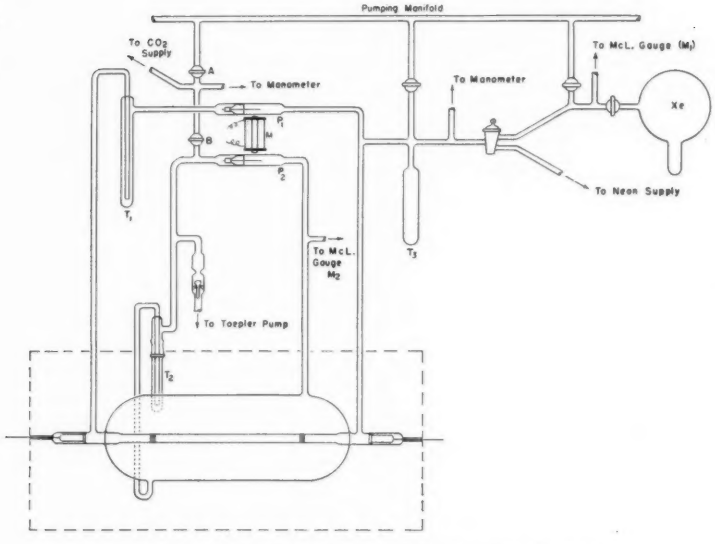

FIG. 2. *Apparatus for photosensitization with Xe* ($^3P_1$).

a volume of 38 cc. while the reaction space has a volume of 2425 cc. When the circulating pump is in operation the gas on leaving the reaction space passes through the trap T_2, containing sodium potassium alloy to adsorb any elemental fluorine produced by the reaction. The alloy is introduced into the trap in an evacuated sealed glass ampoule which is broken when desired by dropping an iron cored glass tube by means of a magnetic hammer.

In making a run the system was first evacuated and the lamp turned on. The lamp was allowed to run until it reached its stable operating state as determined by its appearance and by the current input as measured on a milliammeter. The valve leading to the Toepler pump was closed by raising the mercury in the pump and the desired amount of xenon admitted to the reaction space. The tap B was then closed and the manifold evacuated. The space between A and B was then filled with the desired amount of carbon tetrafluoride and the tap A closed. To start the run the tap B was opened and the pressure allowed to equalize. The tap B was then closed and a pressure measurement taken. The circulating pump was turned on and the ampoule containing the sodium potassium alloy broken.

The run was allowed to proceed until a pressure drop of approximately 10% was noted on the manometer. A sample of the reaction mixture was withdrawn through the Toepler pump for analysis with a mass spectrograph. The contents of the trap T_2 was removed and analyzed for fluorine by the thorium nitrate – alizarin method (3).

III. Experimental Results

A run was made with the Hg (6^3P_1) apparatus without any light, but with the cell at 300° C. and the mercury pump operating. Carbon tetrafluoride was unaffected after 28 hr. under these conditions. The results obtained in subsequent runs are shown in Table I.

TABLE I

PHOTOSENSITIZATION EXPERIMENTS WITH Hg (6^3P_1) RADIATION

Run No.	Cell filling	Reaction time, hr.	Reaction temp., ° C.	Results
1	CF_4: 12.0 mm.	175	24	No pressure change
2	CF_4: 10.7 mm. H_2: 3.6 mm.	205	26	No pressure change
3	CF_4: 9.6 mm.	100	300	No pressure change
4	CF_4: 11.6 mm. H_2: 12.3 mm.	100	300	No pressure change
5	CF_4: 10.65 mm.	25	300	No pressure change

For Runs 1–4 in Table I, pressures were measured with an accuracy of 1 to 2% with a mercury manometer. For Run 5, the pressures were read with an accuracy of 0.1% on a McLeod gauge.

Emeléus *et al.* have reported the existence of compounds of the type CF_3HgX, where X is a halogen atom (1). It was therefore thought possible that the compound CF_3HgF could form in this reaction system. To detect the existence of this type of compound the molecular weight of the gas in the cell was determined before and after each run. No increase in molecular weight was observed. At the end of Run 5, the entire contents of the system was frozen out in a trap with liquid air. The trap was then warmed to −78° C. with dry ice acetone. No product condensable at −78° C. was seen nor was any increase in pressure noted on warming the trap to room temperature. It was therefore concluded that compounds of the type CF_3HgF were not formed.

From the data of Run 5, where the pressures were measured most accurately, it can be calculated that the quantum efficiency for the reaction

$$CF_4 + Hg\ (6^3P_1) \rightarrow CF_3 + Hg\ (^1S_0) + F$$

is less than 3×10^{-4}. That is, the reaction does not proceed to any appreciable extent. It may therefore be concluded that the energy necessary to break the first C–F bond in CF_4 is greater than 112 kcal. per mole.

The results obtained with the 1850 Å lamp are given in Table II.

TABLE II

PHOTOSENSITIZATION EXPERIMENTS WITH Hg (6^1P_1) RADIATION

Run No.	Cell filling	Reaction time, hr.	Reaction temp., °C.	Results
6	CF_4: 11.70 mm.	25	29	No pressure change
7	CF_4: 8.00 mm. H_2: 3.00 mm.	26	27	No pressure change

The pressure readings were accurate to 0.1%. From the data of Run 6, the quantum efficiency of the reaction,

$$CF_4 + Hg\ (6^1P_1) \rightarrow CF_3 + Hg\ (^1S_0) + F$$

is less than 2×10^{-3}; thus the energy necessary to break the C–F bond must exceed 154 kcal. per mole—the energy corresponding to 1850 Å radiation.

For the first run made with the xenon resonance lamp, the reaction space contained 10.0 mm. of carbon tetrafluoride but no xenon. Accordingly, any decomposition would be due to photolysis. After 18 hr. a slight brown deposit was observed on the outside of the quartz tube of the resonance lamp. This solid decomposition product was easily removed by washing with nitric acid. It was concluded that carbon tetrafluoride absorbs 1470 Å, leading to its decomposition.

A second run was then made with conditions identical to the first except that 12 microns of xenon were added to the carbon tetrafluoride in the reaction space. In the same time as previously (18 hr), a much denser brown deposit was visible on the tube. From these preliminary runs, it was concluded that carbon tetrafluoride is photolyzed by 1470 Å radiation, but that the xenon photosensitized decomposition is a more efficient process. Subsequent runs were made at much lower pressures and for shorter durations, in order to avoid reducing the intensity of illumination by a deposit on the lamp window. After the low pressure runs, the lamp output was redetermined by decomposing carbon dioxide to see if the transparency of the quartz tube had been affected. No significant difference was noted.

The data shown in Table III were obtained.

TABLE III

PHOTOSENSITIZATION EXPERIMENTS WITH Xe ($^3P_1$) RADIATION

Run No.	Cell filling	Reaction time, min.	Reaction temp., °C.	Pressure decrease, mm. Hg
8	CF_4: 0.875 mm. Xe : 0.035 mm.	100	23	0.079
9	CF_4: 0.893 mm. Xe : 0.012 mm.	100	26	0.085

Mass spectrograph analyses of the reaction products yielded no indication of any gaseous substance except xenon and carbon tetrafluoride. There was no trace of hexafluoroethane although it was expected, and no gaseous tetrafluoroethylene was found. Fluorine (0.3 mgm.) was found in the alkali metal from the trap (T_2, Fig. 2).

Since there were no gaseous reaction products, the drop in pressure during the reaction is a measure of the carbon tetrafluoride removed from the gaseous phase. From the data in Run 8, a pressure drop of 0.079 mm. is caused by the removal of 0.91 mgm. of carbon tetrafluoride. This corresponds to 0.79 mgm. of fluorine, of which 0.3 mgm. was found in the trap, T_2 (Fig. 2).

It may therefore be assumed that all the carbon corresponding to 0.91 mgm. of carbon tetrafluoride (0.13 mgm.) and the remaining fluorine (0.49 mgm.) remain as the solid polymer. The ratio of C:F is 2:5 and, allowing for the large error possible in the determination of the fluorine in the alkali metal trap, this suggests that the solid observed in the reaction space is the polymer, $(C_2F_4)_n$, polytetrafluoroethylene.

The output of the xenon lamp was considered to be predominantly 1470 Å radiation, with an intensity of 1.9×10^{-9} einsteins per second (4). On the basis of this figure, and the pressure drop observed, the quantum efficiency of the decomposition process is calculated to be 0.9.

IV. Discussion

Hg (6^3P_1) and Hg (6^1P_1) with CF_4

In the introduction to this paper, it was pointed out that the strength of the first C–F bond in CF_4 was probably greater than 130 kcal. per mole. It seemed possible that the bond could be broken by Hg (6^3P_1), since one of the products would be HgF, which has a heat of formation of over 32 kcal. per mole (7). Although the effective cross section of carbon tetrafluoride for quenching the fluorescence of Hg (6^3P_1) is unknown, at a pressure of 1 cm. the collision frequency should be sufficiently high so that virtually all the Hg (6^3P_1) atoms would be quenched. Since no reaction occurred at a measurable rate even at 300° C., it may therefore be concluded that the energy necessary to break the first C–F bond in CF_4 is greater than 112 kcal. per mole.

The Hg (6^1P_1) atoms also failed to produce any decomposition of carbon tetrafluoride. No knowledge of the effective cross section for quenching exists in this case either but since the pressures are sufficiently high, a similar situation exists here as with the Hg (6^3P_1) atoms. Accordingly, it is concluded that the energy necessary to break the first C–F bond in CF_4 exceeds 154 kcal. per mole.

Xe ($^3P_1$) and CF_4

Xenon atoms activated by 1470 Å radiation, with an available energy of 194 kcal. per mole, did decompose the carbon tetrafluoride. The most probable initial step appears to be

$$CF_4 + Xe\ (^3P_1) \rightarrow CF_3 + Xe\ (^1S_0) + F. \qquad (1)$$

No hexafluoroethane was found in the reaction products, indeed no other gaseous carbon fluorides except carbon tetrafluoride. As discussed in the Results, all the fluorine disappearing from the gas phase is accounted for if the solid deposit is the polymer $(C_2F_4)_n$. This seems likely since it would also account for all the carbon contained in the carbon tetrafluoride which disappeared. If we accept Reaction (1) as the initial step, we may conclude that the energy necessary to break the first C–F bond in CF_4 is less than 194 kcal. The route by which CF_3 radicals end up as $(C_2F_4)_n$ is not clear.

An alternative initial step would be

$$CF_4 + Xe\ (^3P_1) \rightarrow CF_2 + F_2 + Xe\ (^1S_0). \qquad (2)$$

The energy available for Reaction (2) is 194 kcal. However, if the F–F bond strength is 55 kcal. Reaction (2) would be energetically possible if the average energy of the first two C–F bonds were 125 kcal. or less. On the basis of the higher value of 70 kcal. for the strength of the F–F bond, the reaction could proceed if the average energy were 132 kcal. or less.

It is interesting to note that since the quantum yield is unity no chain processes occur. Therefore, reactions of the type

$$CF_4 + F \rightarrow CF_3 + F_2$$
$$\text{and}\quad CF_4 + CF_3 \rightarrow C_2F_6 + F$$

do not occur under these conditions.

Hydrogen Atoms with CF_4

From the experimental results of the H atoms with CF_4 we may calculate the minimum activation energy for the reaction

$$CF_4 + H \rightarrow CF_3 + HF \qquad (3)$$

Since the quenching cross section of H_2 for Hg (6^3P_1) is undoubtedly greater than that of CF_4 we estimate that the rate of production of H atoms in Run 5 to be 4×10^{-8} moles per sec. We know that the hydrogen atoms are not affected by collisions with the carbon tetrafluoride molecules and therefore disappear by recombination on the walls. There is no exact knowledge of the life of the hydrogen atoms under these conditions and therefore of their stationary concentration. By estimation from gaseous diffusion rates, assuming the collision diameter of the hydrogen atom to be 4.4×10^{-8} cm., and considering the size and shape of the apparatus, a life of not less than 10^{-2} sec. is considered reasonable. Accepting this estimate and using the usual assumed value of 0.1 for the steric factor it has been calculated from the data of Run 5 that the minimum activation energy for Reaction (3) is 17 kcal. per mole. The steric factor of 0.1 may be too high. It has been reported that the steric factor for the reaction

$$CCl_4 + H \rightarrow HCl + CCl_3$$

is 0.007 (12). Using this lower value and the above assumptions the minimum activation energy for Reaction (3) is 14 kcal. per mole.

Acknowledgment

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DIELECTRIC DISPERSION STUDIES OF POLYVINYL ACETATE IN TOLUENE[1]

By B. L. Funt[2] and S. G. Mason

Abstract

Dielectric measurements were performed on solutions of polyvinyl acetate in toluene at three fixed frequencies over a temperature range. Both the molecular weight of the samples and their concentration in solution were systematically varied. The experiments show that the dielectric behavior of polyvinyl acetate in solution is in better qualitative agreement with the Debye theory of dielectrics than that generally found for solid polymers. Nevertheless serious discrepancies appear when a quantitative comparison is made. It was found that only rough approximations of the molecular size can be made from the Debye theory. Activation energies and entropies for dipole rotation were calculated on the basis of the theory of absolute reaction rates. Small positive values of the entropy were found, indicating restricted dipole rotation. A relatively narrow distribution of relaxation times, approaching that commonly associated with nonpolymeric materials, was found.

Introduction

Measurements of the dielectric constant in the dispersion region can, in principle, yield valuable information regarding molecular properties. Such measurements have been performed on polymeric materials, particularly by Fuoss and coworkers (5) in their extensive investigation of the dielectric properties of various polymers in the solid state.

Little information is available in the literature on the dielectric behavior of polar polymers in solution. A systematic study of the dielectric dispersion of polyvinyl acetate in toluene was undertaken in order to evaluate the applicability of the Debye theory (2) to such a system and to obtain information regarding the dielectric behavior of macromolecules in solution.

For a polar substance a range of frequencies exists within which the dielectric constant drops from its static value at low frequencies to its limiting value at high frequencies. The high frequency value is approximately equal to the square of the refractive index of the substance. In the dispersion region the effective dielectric constant can be considered as a complex quantity given in vector notation as

$$\epsilon^* = \epsilon' - j\epsilon''. \tag{1}$$

The capacitative component ϵ' is the dielectric constant and is normally taken to be the ratio of the parallel geometric capacitance of a condenser containing the dielectric to the capacitance measured *in vacuo*. The dielectric loss factor, ϵ'', appears as an equivalent conductance.

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[2] *At the time holder of a Fellowship under the National Research Council of Canada. Present address: University of Manitoba, Winnipeg, Man.*

The dielectric loss is markedly dependent on the frequency, and the frequency at which the dielectric loss at a given temperature is a maximum may be taken to define a temperature-dependent relaxation time

$$\tau = \frac{1}{\omega_{max}} \quad . \tag{2}$$

In the classical Debye treatment (2) of a rigid dipole undergoing free rotations with linear damping in the alternating electric field, the frequency dependence of ϵ^* at constant temperature is defined by the equations

$$\epsilon' = \epsilon_\infty + \frac{\epsilon_0 - \epsilon_\infty}{1 + \omega^2 \tau^2} \tag{3}$$

$$\epsilon'' = \frac{\epsilon_0 - \epsilon_\infty}{1 + \omega^2 \tau^2} \omega\tau. \tag{4}$$

Here ϵ_0 and ϵ_∞ are the dielectric constants at very low and very high frequencies relative to the dispersion region, and ω is the angular frequency.

Equation (4) predicts a sharply defined maximum of ϵ'' when $\omega\tau = 1$. If the dipole is considered as a sphere of radius a rotating in a medium of viscosity η

$$\tau = \left(\frac{\epsilon_\infty + 2}{\epsilon_0 + 2}\right) \frac{4 \pi \eta a^3}{kT}, \tag{5}$$

where k is the Boltzmann constant and T the absolute temperature. Using this relation an estimate of the size of the rotating unit can be obtained from the relaxation time.

Although these relations provide a reasonable qualitative description of the dielectric behavior of pure liquids and of solutions, the proportionality between τ and η/T applies only approximately as a result of obvious deficiencies in the assumptions made in deriving the relation.

An alternative and more generalized approach in which dipole orientation is treated as a molecular rate process was made by Frank (4) and developed by Eyring (9) and Kauzmann (7). Instead of free rotation restricted only by the steady viscous drag of the medium, dipole orientation is treated as an activated rate process.

Following the principles used in calculating absolute reaction rates, Eyring defines (9) a free energy of dielectric relaxation ΔF^* by the equation

$$\frac{1}{\tau} = \frac{kT}{h} e^{-\Delta F^*/RT} , \tag{6}$$

where h is the Planck constant. This can be rewritten in the form

$$\frac{1}{\tau} = \frac{kT}{h} e^{\frac{\Delta S^*}{R}} e^{-\Delta H^*/RT} , \tag{7}$$

where ΔS^* and ΔH^* are the corresponding entropy and enthalpy of relaxation. It will be noted that the two mechanisms are equivalent when the ΔF^* and

the free energy of activation of the macroscopic fluid for viscous flow are equal (10). This identity has been observed for the lower alcohols from methyl to hexyl alcohol (9), but otherwise fails to apply (10).

Experimental

Apparatus and Procedure

The apparatus and procedure were in many respects similar to those described by Elliott, Jones, and Lockhart (3).

A General Radio Company Type 821-A Twin-T impedance measuring circuit was employed for the radio-frequency measurements. A high frequency generator and a method of null detection were needed in conjunction with the Twin-T. A Clough–Brengle oscillator variable from 0.1 to 40 Mc. was used as a source of high frequency signal. The signal strength from this unit was increased from 0.1 v. to 10 v. by a separate amplifier, and was then fed to the Twin-T network. A Hallicrafter S-140 radio receiver with internal beat oscillator was employed as a highly selective null detector. All units were fed from a constant voltage transformer for increased stability.

Thorough shielding of the apparatus was essential. Coaxial cable, with the outer sheath grounded, was used to connect all units. The ends of the cable were fitted with special jacks for direct shielded connection to all parts of the apparatus. The Twin-T and auxiliary equipment were placed on a copper sheet, and all leads to ground were soldered to it. This sheet was in turn grounded.

In practice an unmodulated signal was fed from the oscillator to the Twin-T network. When the network was unbalanced, part of the signal was fed into the receiver. An internal beat oscillator in the receiver was employed to produce beats with this signal, and these could be heard in the loud-speaker.

When appropriate corrections were applied for residual network impedances at high frequencies, capacitance measurements could be made with an accuracy of $\pm 0.5\%$ and conductance measurements with an accuracy of $\pm 2\%$. The maximum capacitance that could be measured was 1000 $\mu\mu$f. The conductance range was proportional to frequency, being 100 μmhos at 1 Mc. Capacitance changes of 0.1 $\mu\mu$f. could be detected.

In addition to the corrections for residual impedance in the Twin-T circuit, the inductance of the leads of the measuring condenser had to be considered. Although the leads were short (about 3 in.) they could introduce large errors unless corrections were applied. The apparent capacitance C' measured with leads of inductance L_1 is related to the true capacitance C_{xp} as follows

$$C' = \frac{C_{xp}}{1 - \omega^2 L_1 C_{xp}} .$$

The inductance L_1 was evaluated as 1.12×10^{-7} h., and graphical corrections were applied to all data.

Dielectric Cell

A dielectric cell suitable for measurements over a wide temperature range was required. A comparatively simple design shown in Fig. 1 proved to be quite satisfactory. It consisted of two coaxial brass cylinders *E*, *F*, held in position by accurately machined bakelite caps *C*, *G*. The tightly fitting bottom cap, *G*, was sealed with sodium silicate. The centering system employed had the advantage that bakelite was not present between the condenser plates. A ½ in. dead space, *D*, was provided above the inner cylinder to eliminate variations due to incomplete filling or volume shrinkage at low temperatures.

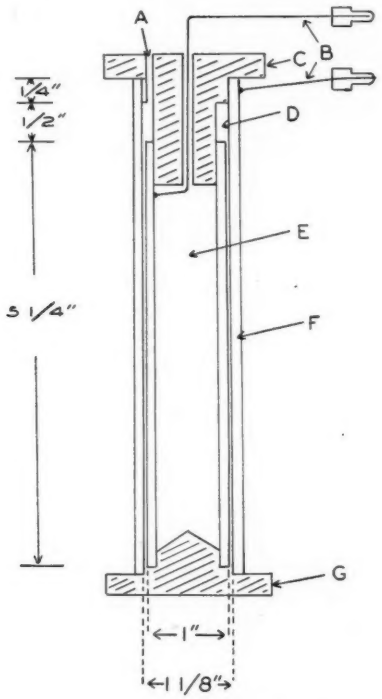

FIG. 1. *Dielectric cell.*

Two short leads of No. 18 B.B.S. wire were soldered directly to the inner and outer cylinders, and were provided with plugs for direct connection to the Twin-T.

The temperature was controlled in a Dewar flask provided with a small heater controlled by a variable transformer. Toluene in the Dewar flask was cooled to approximately −80° C. with dry ice. The cell was then placed in the Dewar flask and allowed to come to thermal equilibrium. Readings were then taken at regular intervals at successively higher temperatures over the

dispersion region. For each reading a temperature rise of 2° per minute was provided for five minutes, and an additional five minutes was allowed for the cell to reach thermal equilibrium. Temperatures were measured on a standardized toluene thermometer immersed in the Dewar flask.

Materials

Samples of polyvinyl acetate were provided through courtesy of the Canadian Resins and Chemicals Co. Three samples of polyvinyl acetate Gelva V 1½, V 15, and V 25 were used. These corresponded to molecular weights of 15,000, 60,000, and 100,000, respectively, as determined by the osmotic pressure method. Reagent grade toluene was dried over sodium metal and distilled prior to use.

The following samples were prepared:

Gelva V 1½:	15, 30, 40, and 60%
Gelva V 15:	15, 30, and 40%
Gelva V 25:	15 and 30%

All concentrations are expressed as grams Gelva per 100 ml. toluene.

Results

Although dielectric dispersion data obtained at fixed temperature and varying frequency are easier to interpret theoretically, most workers in this field have performed their experiments at fixed frequency and varying temperature. The experimental advantages of this procedure are obvious. Measurements performed over a wide temperature range can be used to bring dispersion regions normally out of the scope of a given impedance measuring circuit into its most accurate range. Furthermore, the effects of residual impedances seriously limited the performance of the Twin-T at its higher frequency limits. In fact most of the measurements would have been above the highest frequency limit of the Twin-T at room temperature. No such drawbacks existed at fixed frequencies for temperatures down to −80° C.

Dielectric measurements were made on samples of three molecular weights at concentrations of 15, 30, and 40 gm. per 100 ml. over a temperature range of −80° to 60° C. The corrected values of the equivalent parallel capacitance, C_{xp}, and conductance, G, were interpreted in terms of the components of the complex dielectric constant.

The dielectric cell was calibrated with benzene at 25.0° C., assuming a dielectric constant of 2.273 (3). The geometric capacitance C_g was found to be 54.0 $\mu\mu$f., while the extraneous capacitance C_0 including leads was 29.7 $\mu\mu$f. For measured values of C_{xp} and G, the capacitance and conductance of a sample ϵ' and ϵ'' were evaluated as follows:

$$\epsilon' = \frac{C_{xp} - C_0}{C_g} \tag{8}$$

$$\epsilon'' = \epsilon' \frac{G}{\omega(C_{xp} - C_0)} = \frac{G}{\omega C_g} \quad . \tag{9}$$

At any one frequency, therefore, ϵ'' is proportional to the conductance. The proportionality constant $\frac{1}{\omega C_g}$ was evaluated for 1, 3, and 10 Mc. and the conductance readings were then readily transformed into equivalent values of the dielectric loss.

The relaxation times were evaluated directly from the maxima of the dielectric loss curves at constant frequency. The values obtained in this way were found to be unchanged within experimental error by application of the corrections suggested by Schallamach (10) for the temperature dependence of the relaxation time.

In addition the factor $\frac{\epsilon_0 + 2}{\epsilon_\infty + 2}$ occurring in Debye's equations for relaxation time was set equal to unity. This simplifying assumption is consistent with the recent view of a more linear relation between the dielectric constant and the polarization than that predicted by the Clausius–Mosotti relation (8, 9).

The general nature of the results obtained is shown in Fig. 2 for a typical experiment on 60% Gelva V 1½. It is seen that the dielectric loss curve is

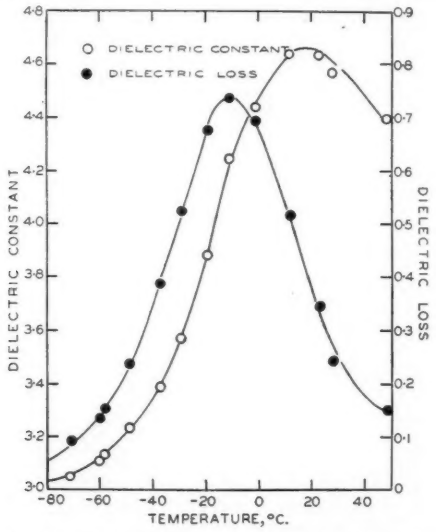

FIG. 2. *Variation of dielectric constant and loss factor with temperature of 60% Gelva V 1½ at 3 Mc.*

symmetrical about a mean temperature, that the position of the maximum is well defined, and that a simultaneous change in dielectric constant occurs in the absorption region. A negative slope in the dielectric constant curve is found at temperatures above those of the absorption region. This is due to the classical variation of dipole polarization with temperature at low frequencies, and follows the relation $P = a + b/T$.

In addition there is a small density effect, but this is almost negligible in comparison to the above, even for the large density temperature coefficient of toluene. This can be verified by observing that at the very low temperature region the effect of this negative slope in dielectric constant is not shown, since there the effect of dipole orientation is absent and the distortion polarization is not temperature dependent. The height of the absorption curve cannot be analyzed directly, since a solution is being investigated and the complex dielectric constant depends on the mole fractions of the components.

It is not possible to discuss in detail the 27 runs performed. The general nature of the data is shown in Fig. 3 where data at one molecular weight

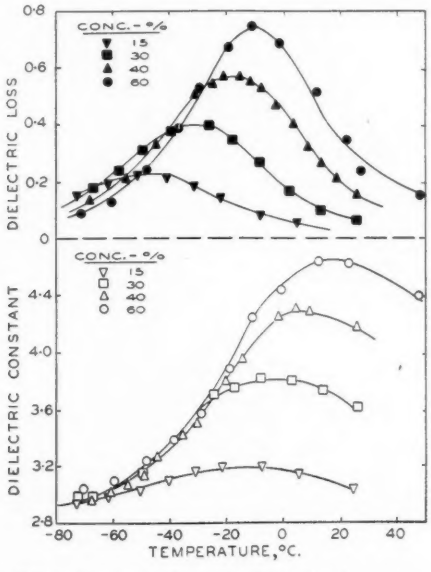

FIG. 3. *Effect of concentration upon the dielectric constant and loss factor of Gelva V 1½ at 3 Mc.*

(V 1½) and one frequency (3 Mc.) are presented at four concentrations 15, 30, 40, 60%. It is seen that the absorption maximum shifts to higher temperatures with increasing concentration. At low temperatures the dielectric constants approach a common value, as is seen in the lower curves. This value, approximately 2.9, is greater than the square of the refractive index of toluene, approximately 2.3. The difference is probably due to the contributions of the dipole polarization of toluene and the atomic polarization, and does not necessarily indicate the presence of another absorption region.

The nature of the results obtained at varying frequency is shown in Fig. 4. Here the results of three experiments at three frequencies (1, 3, 10 Mc.) at fixed concentration (30%) and fixed molecular weight (V 15) are shown.

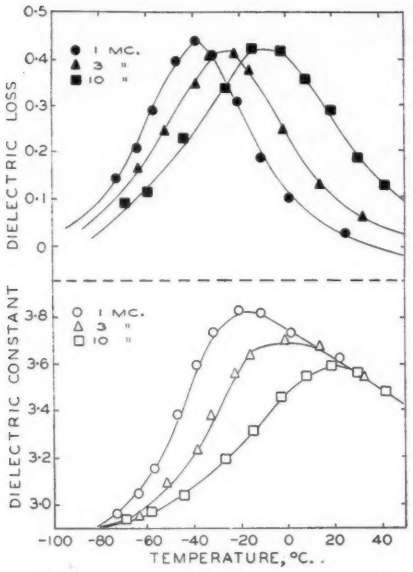

FIG. 4. *Variation of dielectric constant and loss factor with frequency and temperature of 30% Gelva V 15.*

These results are typical of the eight sets of curves obtained in this way. Without exception the absorption maximum shifts to higher temperatures with increasing frequency. At low and high temperatures the dielectric loss approaches zero and the dielectric constants reach common values. The breadth of the absorption curves seems to increase with increasing frequency. This phenomenon is not expected on the basis of the simple Debye theory, but will be considered at a later point in relation to evidence for a distribution of relaxation times. Otherwise the curves seem to show good qualitative agreement with theory.

Discussion of Results

Size of Rotating Units

An approximate size of the rotating unit can be obtained from Equation (5). The viscosity of toluene at various temperatures was obtained by extrapolation of a plot of the data from the International Critical Tables. The radius, a', of the equivalent spherical dipole, and its volume, $V' = \frac{4}{3}\pi a^3$, were thus calculated. These were compared with the equivalent spherical radius and volume of the molecule, a_{op} and V_{op}, calculated from the molecular weight as determined by the osmotic pressure method. The volume, $V_{op} = \frac{M}{\rho N}$, where M is the molecular weight, ρ the density of the solid polymer, and N Avogadro's number. The bulk density of polyvinyl acetate is 1.2.

The data are summarized in Table I.

TABLE I

APPARENT RADIUS AND VOLUME OF ROTATING UNITS

Type Gelva	From Debye dielectric relaxation equations							From osmotic pressure	
		1 Mc.		3 Mc.		10 Mc.			
	Conc., gm./100 ml.	a' $\times 10^8$ cm.	V' $\times 10^{20}$ cm.^{3}	a' $\times 10^8$ cm.	V' $\times 10^{20}$ cm.^{3}	a' $\times 10^8$ cm.	V' $\times 10^{20}$ cm.^{3}	a_{op} $\times 10^8$ cm.	V_{op} $\times 10^{20}$ cm.
V 1½	15	25.1	6.6	19.2	2.94	15.3	1.50	17.1	2.1
	30	30.0	11.2	22.4	4.67	17.1	2.09		
	40	31.9	13.6	24.2	5.90	17.9	2.41		
	60			25.5	7.00				
V 15	15	28.5	9.7	21.1	3.91	16.8	1.96	27.2	8.4
	30	30.4	11.8	23.0	5.10	17.3	2.17		
	40	31.1	12.7	23.6	5.50	17.6	2.31		
V 25	15	28.0	9.2	20.9	3.82	16.0	1.69	32.2	13.9
	30	31.4	13.0	22.7	4.90	17.2	2.14		

The radius and volume calculated from dielectric data increase as the temperature of the absorption maximum decreases. This can be attributed to the unequal temperature dependence of the relaxation time τ and the viscosity η. An equivalent statement is that the activation energies for viscous flow and dipole rotation are not equal. This appears to be generally true, and has been reported by other workers (9).

Extrapolation of the calculated volumes V' in Table I to zero concentration of dissolved polymer shows that the size of the rotating unit is smaller than V_{op}, the volume obtained from the assumed molecular weights determined by the osmotic pressure method. This was not found for the extreme case of the lowest molecular weight (V 1½) at the lowest frequency (1 Mc.). Otherwise, one would conclude that the rotating unit is smaller than the molecule. This is supported by the observation that the increase in size of the rotating units as determined by dielectric measurements at fixed frequency and concentration is much smaller than the expected increase based on the molecular weights of the polymers employed.

The calculated volumes and radii in Table I are large, however, and indicate that the rotating units represent large segments of the polymer molecule.

The limitations of the Debye equation employed in these calculations should not be overlooked. The concept of a dipole as a unit rotating in a medium whose viscosity is the macro-viscosity of the solvent is open to severe criticism (7). In the present work the agreement with the Debye theory appears better than that commonly found in investigations of high polymeric systems (9).

Dipole Moments

A particularly convenient equation, which has been used by several workers for interpreting dielectric dispersion data, has been derived from the Debye theory by Jackson (6).

$$\frac{\epsilon''}{\epsilon'} = \tan \delta = \frac{(\epsilon' + 2)^2}{\epsilon'} \quad \frac{\omega\tau}{1 + \omega^2 \tau^2} \quad \frac{4\pi\mu^2 Nc}{27\ kT},$$

where ϵ' is the dielectric constant of the solution, c is the concentration in moles per milliliter, and the other symbols have their usual significance.

When the absorption maximum is reached $\omega \tau = 1$ and

$$\frac{\omega \tau}{1 + \omega^2 \tau^2} = \frac{1}{2}.$$

The dipole moment can thus be evaluated from the observed maxima of the dielectric loss curve and the corresponding values of the dielectric constant.

In the previous evaluation of the relaxation time from the maxima of the loss curves it was assumed that the experimental value obtained at varying temperature in this way was the same as that obtained at this temperature at varying frequency.

Whiffen and Thompson (12) have suggested that since $\frac{(\epsilon' + 2)^2}{\epsilon'}$ is reasonably constant near the maximum, a plot of $\frac{T \tan \delta}{c}$ versus T would give a plot whose maximum would have the value

$$\frac{1}{2} \frac{(\epsilon' + 2)^2}{\epsilon'} \quad \frac{4\pi\ \mu^2 N}{27k}$$

and thus μ can be evaluated, and any temperature effect compensated.

The data for 15, 30, and 40% V 15 Gelva were plotted in this way in order to evaluate the dipole moment per monomer unit, and to test the method of evaluating the relaxation time. The data are shown in Table II, where values

TABLE II

EVALUATION OF DIPOLE MOMENT PER MONOMER UNIT FOR V 15 GÉLVA AT 3 MC.

Conc., gm./100 ml.	T, °A.	ϵ''	ϵ'	$\tan \delta$	$T \tan \delta$, °A.	T'_{max}, °A.	T_{max}, °A.	μ Debyes
40	235	0.382	3.30	0.115	27.0			
	246	0.469	3.52	0.133	32.8			
	255	0.463	3.76	0.125	31.8			
	265	0.378	3.84	0.099	26.3	250	250	0.91
30	234	0.340	3.23	0.107	24.0			
	241	0.403	3.38	0.119	28.7			
	251	0.408	3.56	0.115	28.8			
	257	0.374	3.64	0.103	26.4			
	272	0.249	3.70	0.067	18.3	247	246	0.98
15	220	0.207	3.02	0.0685	15.1			
	231	0.260	3.13	0.0830	19.2			
	240	0.273	3.22	0.0845	20.3			
	251	0.236	3.30	0.0715	17.9	240	240	1.26

near the dispersion region only have been recorded. The values of T'_{max} were obtained from this plot and should be compared with the values of T_{max} obtained directly from the maxima of the loss curves. The last column gives the calculated values of the dipole moment μ.

The relaxation time constitutes a temperature dependent property of a dielectric and can be considered to define a molecular rate constant. On the basis of the theory of absolute reaction rates a free energy of dipole rotation ΔF^* has been defined in Equation (6).

Differentiation of Equation (7) yields the result

$$\Delta H^* = R(d \ln \tau / d\, 1/T) - RT .$$

If σ is the slope of the curve when log τ is plotted against $1/T$,

$$\Delta H^* = 2.303\, R\sigma - RT .$$

Plots of log τ versus $1/T$ yielded straight lines and thus permitted an evaluation of ΔH^*.

Values of ΔS^* were obtained by subtraction of the calculated values of ΔF^* from ΔH^*. The results are summarized in Table III.

TABLE III

EVALUATION OF FREE ENERGY, ENTHALPY, AND ENTROPY

Type Gelva	Conc., gm./100 ml.	1 Mc.		3 Mc.		10 Mc.		ΔH^* kcal.	ΔS^* cal/°A.
		T_{max}, °A.	ΔF^* kcal.	T_{max}, °A.	ΔF^* kcal.	T_{max}, °A.	ΔF^* kcal.		
V 1½	15	213	5.74	223	5.55	246	5.59	6.97	5.9
	30	233	6.30	243	6.09	262	5.95	9.51	14.2
	40	241	6.55	256	6.46	269	6.15	9.91	14.3
	60			263	6.62				
V 15	15	227	6.15	240	5.99	259	5.89	7.59	6.0
	30	235	6.38	246	6.12	264	6.05	8.90	11.2
	40	238	6.48	250	6.26	267	6.09	9.92	10.4
V 25	15	223	6.05	234	5.85	253	5.72	8.38	11.0
	30	239	6.46	245	6.11	263	6.00	11.31	20.5

A high entropy of activation such as the 150 entropy units reported for tetrachlorodiphenyl (7) requires the contributions of several rotating units (9). Since a molecule in the activated state is rotating and has more configurations available to it than a molecule at rest, this increase in rotational entropy is shown in a positive value of ΔS^*. When the entropy change is very large it is believed to represent the contributions of several molecules.

Entropies of zero or small negative values have been found for hexasubstituted benzenes (13) and for highly plasticized polyvinyl chloride. This has been interpreted (9) as indicating that the dipoles are rotating freely in the normal state.

The values of ΔF^*, ΔH^*, and ΔS^* obtained in the present work are approximately the same as those obtained for alcohols of about six to eight chain atoms. The positive entropies indicate that the dipole rotation is restricted. The relatively small magnitude of these entropies, however, shows that the hindrance to rotation is not as great as in many polymers and other systems where higher entropy values are found.

Extrapolation of the plots of log τ vs. $1/T$ for 30% Gelva V 1½, V 15, and V 25 yields the value of $1.5 \pm 2 \times 10^{-9}$ sec. for the relaxation time at 25.0° C.

The data have been analyzed to determine whether a distribution of relaxation times was indicated. A Cole and Cole circular arc plot (1) of ϵ' versus ϵ'' at constant temperature was tried. Three pairs of values of ϵ' and ϵ'' were obtained by proceeding in a vertical direction on curves such as shown in Fig. 4. These were sufficient to define a semicircle, and values of ϵ_0 and ϵ_∞ could be obtained by extrapolation. For 40% V 1½ a value of 0.312 radians was obtained for the Cole and Cole $\alpha(1)$. This can be compared with values of the same magnitude recently obtained by Smyth and coworkers for chlorinated hydrocarbons (11), and with the much wider distributions found for polyvinyl chloride (7).

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THE DIFFUSION OF POLYSTYRENE IN DIFFERENT SOLVENTS[1]

By P. Z. Adelstein[2] and C. A. Winkler

Abstract

The variation of the diffusion coefficient of polystyrene with molecular weight in bromobenzene, toluene, and methyl ethyl ketone, obtained with a modified Northrop cell, obeyed the relation $1/D = AM^{\epsilon}$ where A and ϵ are empirical constants. The decrease in the exponent ϵ with increasing molecular weight agrees qualitatively with the predictions of Debye and Bueche. Both the exponent and the ratio of the diffusion coefficients of a low to high molecular weight sample increased in better solvents, indicating that in such solvents the polystyrene molecules are elongated. Using this ratio as a criterion of solvent power, there is good agreement with the results of swelling and precipitation studies. The decrease in the diffusion coefficient – solvent viscosity product for a low and a high molecular weight polymer in good solvents suggests an increase in the degree of solvation of the polymer. The variation of this product with solvent composition in solvent mixtures indicates the existence of selective solvation by the good solvent. The activation energies of diffusion were of the order of 3 kcal. for both high and low molecular weight polymers in good and poor solvents. This supports Eyring's theory of segment flow for polymer diffusion.

Introduction

It is generally assumed that the rate of diffusion of a polymer in solution is governed largely by the shape of the polymer chain which, however, need not be the same in different solvents. It seems reasonable to regard a thermodynamically favorable, or good, solvent as one in which the polymer chain tends to stretch out so as to permit maximum polymer–solvent contact. An unfavorable or poor solvent would then be one in which there is little polymer–solvent attraction and the polymer molecule tends to coil up and assume a more nearly spherical shape.

Evidence for a change of polymer configuration in different solvents has been obtained indirectly from the variation of the intrinsic viscosity with solvent (1, 13, 18, 24), and directly from the dissymmetry of light scattering through high polymer solutions (7, 26).

The diffusion of high polymers has been given several theoretical treatments (6, 9, 10, 11, 21, 25), and the dependence of the diffusion coefficient on the molecular weight, M, expressed (6, 8, 24) by relations of the form.

$$1/D = AM^{\epsilon}, \tag{1}$$

where A and ϵ are constants for a given polymer–solvent system. The different geometrical models used to represent the polymer molecules lead to different values for the exponent, ϵ, but it is generally agreed that the value of ϵ should increase as the shape of the molecules becomes more elongated, as for example, if a good solvent is substituted for a poor one.

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Contribution from the Department of Physical Chemistry, McGill University, Montreal, Que. with financial assistance from the National Research Council of Canada.

[2] *Holder of Bursary and Studentships under the National Research Council of Canada.*

The various treatments of polymer diffusion referred to above regard the solvent as a continuous medium and the polymer molecule to move as a unit. In contrast, it is also possible to regard the polymer as moving by segments into holes created in the solvent, the size of a hole being that necessary to accommodate a segment (12, 15, 20). The magnitude of the diffusion coefficient of a polymer in solution nevertheless depends on the total chain length, since the probability of co-ordinated motion of all segments decreases as the chain length increases.

The experiments described in the present paper were made principally to obtain information about the way in which changes in rate of diffusion of a high polymer may reflect changes in its shape in different solvents. The data obtained are possibly of some interest also in providing at least qualitative tests of the existing theories of diffusion.

Apparatus

A modified Northrop cell (23) was used in this investigation (Fig. 1). The cell *A*, with a volume of about 35 cc., contained the polymer solution, and was immersed in about 95 cc. of solvent in the outer jacket, *B*. Diffusion of the polymer solute took place through the sintered glass membrane, *C*. In place of the usual stopcock on the Northrop cells, there was a ground glass joint, *D*, into which was fitted a closed plug. Since volatile solvents were used, the outer jacket, *B*, was fitted with a ground glass joint, *E*, and the

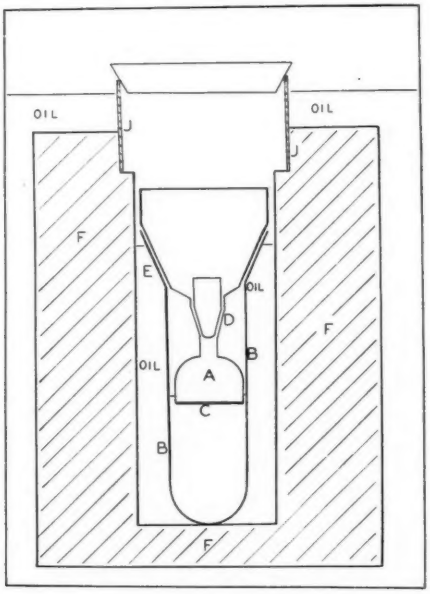

FIG. 1. *Northrop diffusion cell.*

cell was attached to the inner joint. This gave a tight seal to prevent loss of solvent by evaporation. To maintain atmospheric pressure in the jacket at elevated temperatures, a slight groove was made in the glass joint, *E*.

Temperature fluctuations and space variations of temperature were minimized by placing the diffusion cell inside an aluminum cylinder, *F*, the latter being immersed in an oil bath regulated to within $\pm 0.05°$ C. A glass cylinder, *J*, attached to the inner wall of the aluminum cylinder, and protruding above the surface of the oil in the bath, prevented the bath oil from overflowing into the cell.

The experimental technique for using this cell was similar to that already described by Northrop (23) and McBain (22). The amount of solute diffusing through the membrane in a given time was found by evaporating the solvent and weighing the residual polymer.

The diffusion coefficient was calculated, using the integrated form of Fick's First Law of Diffusion,

$$BDt = \log \frac{\triangle C_0}{\triangle C_f}$$

where $\triangle C_0$ is the initial concentration gradient across the membrane, $\triangle C_f$ is the final concentration gradient after time t, and B is the cell constant (17, 22).

Following the procedure of Gordon (17), the value of B was obtained by calibrating the diffusion cells with 0.1 *N* potassium chloride solution diffusing into water at 25.0° C., the value of D being taken as 1.834×10^{-5} cm^2 per sec.

No attempt has been made to correct for possible errors due to stagnant layers adjacent to the diaphragm, which might exist with the density stirring used. However, it seems reasonable that such errors should not affect the main conclusions of the present study, which are based on the relative behavior of polymers of the same material at the same concentration and temperature.

Materials

Polystyrene was chosen as the polymer for this study for two reasons: (i), it is a flexible unbranched chain molecule which might be expected to change shape with solvent; and (ii), a large number of studies have already been made with this polymer by methods such as viscosity and light scattering. With the exception of a commercial "Monsanto" polystyrene, the different polymer samples were prepared by polymerizing styrene at elevated temperatures in benzene, using benzoyl peroxide as catalyst. The polymers were fractionally precipitated with methanol from 5% benzene solution. The samples of polystyrene obtained are shown in Table I.

TABLE I

MOLECULAR WEIGHT OF POLYSTYRENE SAMPLES

Polymer	Molecular weight
A-s	300
A	500
C	1,300
D-2	6,500
E-1-2	20,000
G-2	130,000
"Monsanto"	340,000
Mon.-3	430,000
B	560,000

The molecular weights of the different polymer samples were obtained from viscosity measurements, using the viscosity molecular weight equation of Mark, $[\eta] = KM^a$. The constants K and a were taken as 3.7×10^{-4} and 0.62 respectively, as determined by Goldberg, Hohenstein, and Mark (16).

Results and Discussion

Typical results obtained from the diffusion experiments are shown in Fig. 2 for 1.935% "Monsanto" polystyrene in chlorobenzene at 60.0° C.

The linear relations between amount of material diffused and time indicate that the small amount of diffused polymer did not appreciably alter the

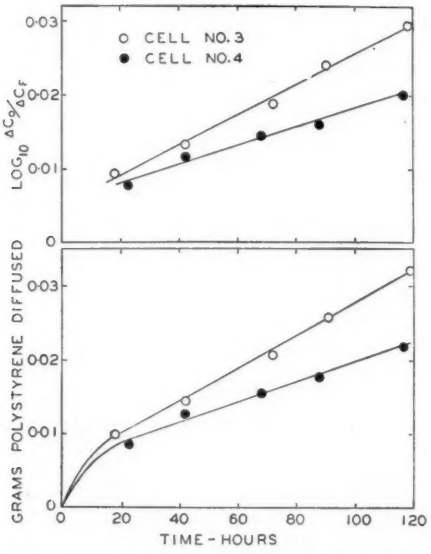

FIG. 2. *Typical diffusion rate curves and log plots.* 1.935% *"Monsanto" Polymer in chlorobenzene at 60.0° C.*

concentration gradient across the membrane. These rate lines, when extrapolated, did not pass through the origin, since at zero time the concentration in the membrane was not that for steady state conditions, but was that of the original solution. However, once diffusion had proceeded for about 20 hr., steady state conditions were obtained and the diffusion rate decreased to its normal value. The straight lines obtained by plotting log $\Delta C_0/\Delta C_f$ against time also do not pass through the origin (Fig. 2) but the values of BD calculated from their slopes should still be essentially correct.

Effect of Molecular Weight

The effect of molecular weight upon the diffusion coefficient was determined in the solvents bromobenzene, toluene, and methyl ethyl ketone, and these results are plotted in Fig. 3. Swelling measurements on cross-linked poly-

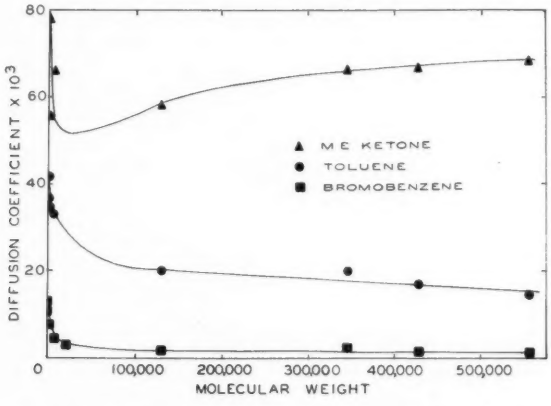

FIG. 3. *Effect of molecular weight upon the diffusion coefficient.* $\Delta C = 1.930$ *(average). Temp.* $= 60.0°$ *C.*

styrene in these solvents show that they may be regarded as examples of a good, indifferent, and poor solvent respectively (4). While the general type of inverse relation between molecular weight and the diffusion coefficient, predicted by many workers, is found to be valid for the solvents bromobenzene and toluene, it is evident that the relation for methyl ethyl ketone is different.

A plot of log $1/D$ against log M for the results obtained in bromobenzene is shown in Fig. 4. From this curve the value of the exponent ϵ, in Equation (1), is found to decrease from 0.37 to 0.20 as the molecular weight increases from 300 to 430,000. A similar plot for the results in toluene shows too much scatter of the points to infer a change in the value of the exponent with molecular weight, but an average value of 0.12 was calculated. With methyl ethyl ketone as solvent, a satisfactory value for the exponent could not be obtained owing to the peculiar variation of the diffusion coefficient with molecular weight (Fig. 3).

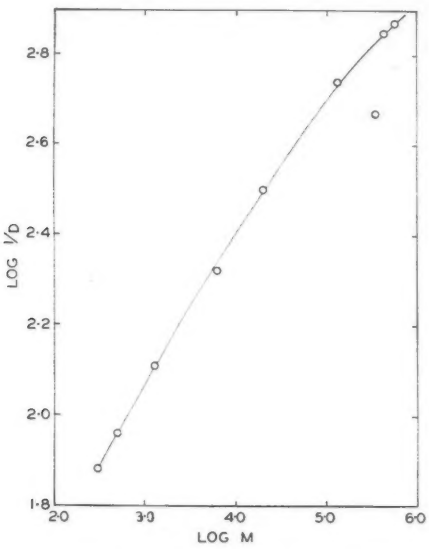

FIG. 4. *Plot of log 1/D against log M. Solvent = Bromobenzene.* $\Delta C_0 = 1.937$. *Temp.* $= 60.0°$ *C.*

The experimental values of ϵ may be compared with those predicted by the various theories, which are 0.5 to 1.0 (6, 8) and 0.33 to 0.78 (24). Also the plot of DZ against $Z^{1/2}$ in Fig. 5 for the results in bromobenzene is to be compared with a predicted linear relation (21). However, these comparisons are of doubtful significance since values of D extrapolated to zero concentration should be used, to eliminate the dependence on concentration. Because of the difficulties involved in making accurate diffusion measurements at low

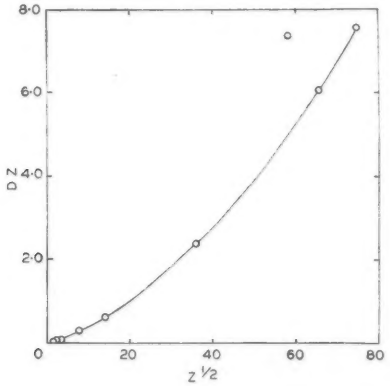

FIG. 5. *Plot of* DZ *against* $Z^{\frac{1}{2}}$. *Solvent = Bromobenzene.* $\Delta C_0 = 1.937$. *Temp.* $= 60.0°$ *C.*

concentrations, this was not done in the present study. Qualitatively the results shown in Fig. 4 agree with the prediction (6) that ϵ decreases as the molecular weight of the polymer increases.

Although the absolute values of ϵ obtained in these experiments probably have little significance, the relative values for the different solvents should give an indication of the relative shapes of the polymer molecules in these solvents. A smaller value of the exponent ϵ in a poor solvent than in a good solvent is associated with a more coiled polymer molecule in the poor solvent (8, 24), a decrease in effective size of the molecule being reflected in a smaller dependence of the diffusion coefficient upon molecular weight. The ϵ values and the diffusion coefficient – molecular weight curves for bromobenzene and toluene are as expected, but the results obtained with methyl ethyl ketone are difficult to interpret. It is possible that because of the strong polymer–polymer attraction in this solvent, the molecular chains form loosely coiled networks, such that there is little decrease in driving force for diffusion but relatively large decreases in resistance to flow compared with that offered if the chains diffused separately as more linear units. If now, as seems reasonable, chains of large molecular weight suffer relatively more coiling in the networks than chains of lower molecular weight, it becomes possible to account for the observed increase in the diffusion coefficient with molecular weight in methyl ethyl ketone. The decrease in diffusion coefficient up to molecular weights of the order of 25,000 would simply imply that, in this range, the molecules are not sufficiently flexible to permit effective coiling, and a normal decrease in diffusion coefficient with increased molecular weight is observed.

Effect of Concentration

It has already been indicated (3) that the diffusion coefficient for a given polymer–solvent system varies with the concentration. Diffusion coefficients obtained at various concentrations in bromobenzene and methyl ethyl ketone for the high molecular weight "Monsanto" polymer and the low molecular weight polymer *C* are summarized in Table II. Since these diffusion coeffi-

TABLE II

EFFECT OF CONCENTRATION ON THE DIFFUSION OF POLYSTYRENE

Temp. = 60.0° C.

Solvent	Concentration, gm./100 cc. soln.	$D \times 10^3$ cm.²/hr.	
		Polymer *C*	"Monsanto" polymer
Bromobenzene	0.969	9.48	2.27
	1.937	7.78	2.16
	2.906	8.45	2.12
Methyl ethyl ketone	0.962	54.7	45.5
	1.962	55.2	66.2
	2.885	106.0	75.0

cients are average or integral values they cannot be used to obtain a quantitative relation between diffusion coefficient and concentration, but they do indicate the general trend.

The slight decrease in the diffusion coefficient with increasing concentration in the good solvent, bromobenzene, would be expected if at low concentrations the relatively elongated polymer molecules move more or less independently of one another, while at higher concentrations they tend to interfere with one another. The marked increase in the diffusion coefficient with concentration in the poor solvent, methyl ethyl ketone, would result if the polymer molecules move independently of one another at low concentrations, but suffer increased intermolecular association (19), so as to decrease the resistance to flow in the manner indicated previously.

An important fact is that the effect of concentration is similar for both the high and low molecular weight samples. This justifies the use previously made in this study of the diffusion coefficients obtained at some arbitrary concentration, rather than the extrapolated values at zero concentration, as an indication of the relative shapes of the polymer molecules in different solvents.

Effect of Solvent

(*a*) *Molecular Shape and Solvent Power*

If the subscripts l and h refer to a low and a high molecular weight polymer, and $\epsilon_{av.}$ is the average value of the exponent for a given solvent, then,

$$\frac{D_l}{D_h} = \frac{A\ M_h^{\epsilon_{av.}}}{A\ M_l^{\epsilon_{av.}}} = \left(\frac{M_h}{M_l}\right)^{\epsilon_{av.}}.$$

Hence, like ϵ itself, the ratio of the diffusion coefficients of a low to high molecular weight polymer can be used as an indication of solvent power. Poor solvents, for which the value of the exponent is small, should have a low value for the ratio of the diffusion coefficients, while in good solvents both the exponent and diffusion coefficient ratio are greater.

The significance of this ratio can be seen if, as a first approximation, a low molecular weight polymer is considered to have the same shape in all solvents. Its diffusion coefficient should depend primarily on the properties of the solvent itself permitting the solvent effect to be studied independently of the shape of the polymer. The diffusion coefficient of a high molecular weight polymer, however, should also reflect the molecular shape of the solute, and its resistance to flow should be greater in solvents where it has an elongated shape. Therefore the ratio of the diffusion coefficients of a low to high molecular weight polymer should be greater in solvents where the polymer chains are extended.

The determination of solvent power by the ratio of the diffusion coefficients of a low to a high molecular weight polymer is a much less laborious procedure than that involved in obtaining the exponent.

The diffusion coefficients of Polymer *A* and "Monsanto" polymer were measured in six solvents, and the ratio of the diffusion coefficients of the low to high molecular weight samples are recorded in Table III.

TABLE III

DIFFUSION COEFFICIENT RATIO AND SOLVENT POWER

ΔC_0 = 1.930 (average)

Solvent	*D* ratio low/high 60.0° C.	Weight Swelling Ratio	$[\eta]$ 60.0° C.	η_G	Mole % heptane	Mole % butanol	Mole % methanol
Methyl ethyl ketone	1.18	5.47	0.49	0.53	35.3	28.1	23.5
Ethyl acetate	1.28	4.93	0.51	0.55	36.0	26.1	20.9
Toluene	1.83	11.06	1.00	0.44	44.2	39.9	43.1
Benzene	2.03	11.27	1.09	0.45	42.0	37.2	39.6
Chlorobenzene	5.00	15.96	1.06	0.42	51.8	42.7	46.6
Bromobenzene	5.09	19.51	1.08	0.43	53.2	43.5	46.7

To compare the diffusion coefficient ratio with other criteria of solvent power, swelling measurements in the same six solvents for polystyrene cross-linked with divinylbenzene are also tabulated. These observations were made by Boyer and Spencer (4), and are reported as weight swelling ratios, defined as the ratio of the weight of the swollen polymer to that of the original polymer. This ratio should be greatest in the good solvents. From their measurements on swelling, these authors calculated the gel–solvent interaction constant μ_G using the Flory–Rehner expression for the activity of the solvent in the gel phase (14). This constant increases in value as the solvent becomes poorer.

A further indication of solvent power is given by the mole per cent of the nonsolvents heptane, butanol, and methanol required for precipitation of "Monsanto" polymer from a 2% (2 gm. per 100 cc. of solution) solution at 20° C. The intrinsic viscosities at 60.0° C. for the same polymer sample in these six solvents are also given.

The solvent powers of the liquids as indicated by precipitation data are found to be practically independent of the nonsolvent used, and agree with the results of swelling measurements. Although the intrinsic viscosities indicate a marked difference between the poor solvents, methyl ethyl ketone and ethyl acetate on one hand, and the indifferent (benzene and toluene) and good (bromobenzene and chlorobenzene) solvents on the other, no difference is indicated between these last two groups.

It is evident that the diffusion coefficient ratios increase as the solvent becomes better, and provide verification of the concept that a polymer chain becomes elongated in good solvents.

(*b*) *Degree of Solvation*

Diffusion experiments, besides reflecting the shape of the polymer solute, can also indicate the extent of solvation of the solute by the variation in the product of the diffusion coefficient and solvent viscosity with change in solvent. The product of the diffusion coefficient, *D*, and the solvent viscosity, η, should remain constant for all solvents, provided the size and

permeability of the molecules remain unchanged (6, 9, 10, 11, 25). This assumption can be made for a very low molecular weight polymer. The values of this product, $D\eta$, for Polymer *A* (mol. wt. = 500) and "Monsanto" polymer (mol. wt. = 340,000) in six solvents are given in Table IV.

TABLE IV

EFFECT OF SOLVENT ON PRODUCT OF DIFFUSION COEFFICIENT AND SOLVENT VISCOSITY

$\Delta C_0 = 1.930$ (average)

Solvent	$D\eta \times 10^3$, cm.²/hr. Polymer *A*	$D\eta \times 10^3$, cm.²/hr. "Monsanto" polymer
Methyl ethyl ketone	22.1	18.8
Ethyl acetate	15.5	12.1
Toluene	13.9	7.58
Benzene	15.8	7.79
Chlorobenzene	6.76	1.35
Bromobenzene	7.75	1.52

Instead of remaining independent of solvent, the $D\eta$ product for polymer *A* was found to decrease as the solvent became better. This behavior can be attributed to solvation of the polymer by the solvent molecules such that the effective radius is increased and the product $D\eta$ decreased. As the solvent becomes poorer, the degree of solvation becomes less and the product $D\eta$ is decreased relatively less.

The decrease in the product $D\eta$ for the high molecular weight sample as the solvent becomes better is apparently the result of both solvation and uncoiling of the molecule. This is indicated by the fact that the product of $D\eta$ for Polymer *A* changed by a factor of 3.3, compared with a change by a factor of 12.4 for "Monsanto" polymer.

(*c*) *Solvent Mixtures*

Since both the polymer shape and the degree of solvation would be expected to decrease as a nonsolvent is added to a polymer solution, diffusion measurements were made in bromobenzene–butanol mixtures to determine its behavior. In Fig. 6, $D\eta$ is plotted as a function of the mole per cent bromobenzene for Polymer *C* and "Monsanto" polymer.

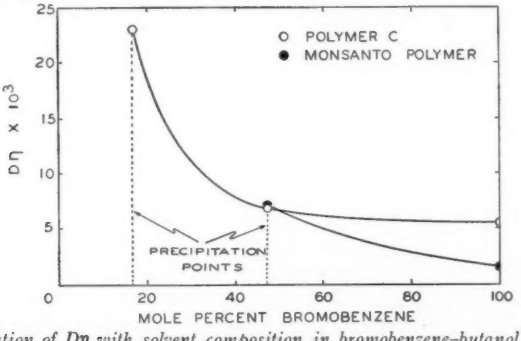

FIG. 6. *Variation of $D\eta$ with solvent composition in bromobenzene–butanol mixtures.* $\Delta C_0 = 1.937$. *Temp.* = 60.0° C.

If the molecular shape of Polymer *C* (mol. wt. = 1300) can be considered to be essentially independent of solvent, the variation in $D\eta$ for this polymer should largely reflect the change in solvation. Fig. 6 shows that over most of the composition range, the value of $D\eta$ is closer to its value in bromobenzene than would be expected. This indicates that the polymer molecule in a mixed solvent is preferentially surrounded by the better solvent. Because of selective solvation, changes in solvation and molecular shape should be small, but the variation in $D\eta$ for "Monsanto" polymer should be greater than that for Polymer *C* over the same concentration range. This was found experimentally.

The variation of $D\eta$ for the same polymers with the composition of a bromobenzene – methyl ethyl ketone mixture is shown in Fig. 7. The evidence

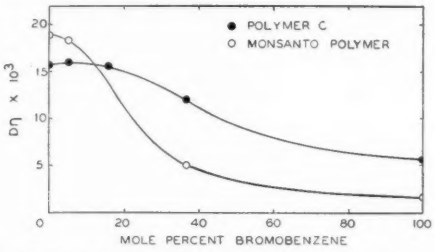

FIG. 7. *Variation of $D\eta$ with solvent composition in bromobenzene – methyl ethyl ketone mixtures. $\Delta C_0 = 1.925$. Temp. = 60.0° C.*

for selective solvation by the bromobenzene is not pronounced, but there is definite indication that the degree of solvation and shape of a polymer chain is only slightly affected by methyl ethyl ketone addition until there is insufficient bromobenzene present for complete solvation of the polymer.

The variation in the ratio of the diffusion coefficients of Polymer *C* and "Monsanto" polymer with solvent composition (Fig. 8) seems to indicate that the shape of "Monsanto" polymer changed only slowly with a decreasing concentration of bromobenzene, until the solvent mixture was 50 mole% of bromobenzene. The polymer then rapidly assumed the coiled configuration which it has in the poor solvent, methyl ethyl ketone.

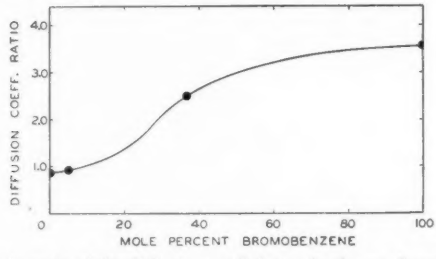

FIG. 8. *Variation in ratio of the diffusion coefficients of polymer C and "Monsanto" polymer with solvent composition in bromobenzene – methyl ethyl ketone mixtures. $\Delta C_0 = 1.925$. Temp. = 60.0° C.*

Effect of Temperature

The activation energies of diffusion were obtained for Polymer *C* and "Monsanto" polymer in the good solvent bromobenzene and the poor solvent methyl ethyl ketone. These results are recorded in Table V.

TABLE V

ACTIVATION ENERGIES OF DIFFUSION

$\Delta C_0 = 1.930$ (average)

Polymer	Activation energy, kcal./mole	
	Bromobenzene	Methyl ethyl ketone
C	4.0	2.1
"Monsanto"	3.5	2.4

The difference between the activation energies for the high and low molecular weight polymers is less than the experimental error, and the activation energies in both solvents are of the same order of magnitude. These results suggest that the polymer moves as a series of segments, in agreement with Eyring's treatment.

Since Eyring considered the diffusion of large solute molecules to be governed by the movement of the solvent molecules, the activation energy of diffusion of the polymer should be equal to that for viscous flow of the solvent. The product of the diffusion coefficient and solvent viscosity should then be independent of temperature. The $D\eta$ values for both polymers in bromobenzene and methyl ethyl ketone at four temperatures are given in Table VI.

TABLE VI

EFFECT OF TEMPERATURE ON THE PRODUCT OF DIFFUSION COEFFICIENT AND SOLVENT VISCOSITY

Solvent	Temp., ° C.	$D\eta \times 10^3$		Ratio of diffusion coefficient, *C*/"Monsanto"
		Polymer *C*	"Monsanto" polymer	
Bromobenzene ($\Delta C_0 = 1.937$)	69.7	7.50	1.61	4.67
	60.0	5.48	1.52	3.60
	44.8	8.20	1.29	6.30
	29.8	4.96	1.26	3.90
Methyl ethyl ketone ($\Delta C_0 = 1.925$)	69.7	15.7	19.9	0.79
	60.0	15.7	18.8	0.83
	44.8	16.7	18.5	0.90
	29.8	15.3	18.3	0.83

The value of $D\eta$ decreased considerably with temperature in bromobenzene, but was practically constant in methyl ethyl ketone. This is an indication that at lower temperatures bromobenzene becomes a better solvent, while the solvent power of methyl ethyl ketone remains essentially unaltered.

Increase in solvent power of a good solvent as the temperature is decreased has been observed also by viscosity (2, 5) and light scattering (26) measurements.

The ratio of the diffusion coefficients of Polymer *C* to "Monsanto" polymer are also given in Table VI. Because of the uncertainty of several values of the diffusion coefficient fo· Polymer *C* in bromobenzene, no information can be obtained from the ratios in this solvent. With methyl ethyl ketone, these ratios do not indicate any appreciable change in shape of the polymer with temperature.

Acknowledgments

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SYNTHESIS OF AMINO ACIDS FROM SUBSTITUTED CYANOACETIC ESTERS[1]

By Paul E. Gagnon, Jean L. Boivin[2], and Paul A. Boivin[3]

Abstract

Nine α-amino acids, namely, *dl*-α-aminoundecylic acid, *dl*-α-aminostearic acid, *dl*-α-amino-β-methylcaproic acid, *dl*-α-amino-β-ethylvaleric acid, *dl*-α-amino-β-methylenanthic acid, *dl*-α-amino-α-acetylacetic acid, *dl*-α-amino-γ-cyclohexylbutyric acid, *dl*-α-(1-indanyl)-α-aminoacetic acid, and *dl*-proline have been prepared from ethyl alkylcyanoacetates. When a mixture of formic and hydrochloric acids and water was used, instead of hydrochloric acid, as hydrolyzing agent, α-amino-β-ethylvaleric acid and α-(1-indanyl)-α-aminoacetic acid were obtained with better yields. The following compounds, as far as the authors are aware, have been prepared for the first time: ethyl α-cyanoundecylate, α-cyanoundecylic hydrazide, anisal α-cyanoundecylic hydrazide, α-cyanoundecylic azide, α-carbethoxyaminoundecylonitrile, *dl*-α-aminoundecylic acid, 5-(*n*-nonyl)-hydantoin, α-cyanostearic hydrazide, anisal α-cyanostearic hydrazide, α-cyanostearic azide, α-carbethoxyaminostearonitrile, α-cyano-β-methylcaproic hydrazide, anisal α-cyano-β-methylcaproic hydrazide, α-cyano-β-methylcaproic azide, α-carbethoxyamino-β-methylcaproonitrile, *dl*-α-amino-β-methylcaproic acid and its copper salt, α-cyano-β-ethylvaleric hydrazide, anisal α-cyano-β-ethylvaleric hydrazide, α-cyano-β-ethylvaleric azide, α-carbethoxyamino-β-ethylvaleronitrile, *dl*-α-amino-β-ethylvaleric acid, and its copper salt, ethyl α-cyano-β-methylenanthate, α-cyano-β-methylenanthic hydrazide, anisal α-cyano-β-methylenanthic hydrazide, α-cyano-β-methylenanthic azide, α-carbethoxyamino-β-methylenanthonitrile, *dl*-α-amino-β-methylenanthic acid, and its copper salt, α-cyano-α-acetylacetic hydrazone, anisal α-cyano-α-acetylacetic hydrazone, α-cyano-α-acetylacetic azide, α-carbethoxyamino-α-acetylacetonitrile, *dl*-α-amino-α-acetylacetic acid, ethyl-α-cyano-γ-cyclohexylbutyrate, α-cyano-γ-cyclohexylbutyric hydrazide, anisal α-cyano-γ-cyclohexylbutyric hydrazide, α-cyano-γ-cyclohexylbutyric azide, α-carbethoxyamino-γ-cyclohexylbutyronitrile, 5-(β-cyclohexylethyl)-hydantoin, 1-indanyl bromide, ethyl α-(1-indanyl)-α-cyanoacetate, α-(1-indanyl)-α-cyanoacetic hydrazide, anisal α-(1-indanyl)-α-cyanoacetic hydrazide, α-(1-indanyl)-α-cyanoacetic azide, α-(1-indanyl)-α-carbethoxyaminoacetonitrile, *dl*-α-(1-indanyl)-α-aminoacetic acid, and its copper salt, ethyl δ-chloro-α-cyanovalerate, δ-chloro-α-cyanovaleric hydrazide, anisal δ-chloro-α-cyanovaleric hydrazide, δ-chloro-α-cyanovaleric azide, δ-chloro-α-carbethoxyaminovaleronitrile.

Introduction

The Curtius reaction is a preparative method for isocyanates and their derivatives, such as urethanes, ureas, amides, and amines. The first application of this degradation for the preparation of an amino acid from cyanoacetic ester was suggested in 1915 by Darapsky and Hillers (4), who obtained glycine.

In 1936, the same method was used to prepare α-amino-*n*-valeric acid, α-aminoisoamylacetic acid, and *dl*-leucine (3) from the corresponding alkylcyanoacetic esters. Gagnon, Gaudry, and King (6), in 1944, successfully synthesized α-amino-δ-phenoxyvaleric acid, *dl*-methyltyrosine, *dl*-phenylalanine, *dl*-valine and, in 1947, Gagnon, Savard, Gaudry, and Richardson (8) obtained *dl*-norleucine, *dl*-isoleucine, *dl*-S-benzylhomocysteine, and *dl*-methionine. Gagnon and Boivin (5), in 1948, prepared nine α-amino acids,

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Contribution from the Department of Chemistry, Laval University, Quebec. This paper constitutes a part of a thesis submitted to the Graduate School, Laval University, in partial fulfillment of the requirements for the degree of Doctor of Science.

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[3]*Holder of a Canadian Industries Limited Research Scholarship in 1948–1949. Holder of a Bursary of the Scientific Research Bureau of the Province of Quebec in 1949–1950.*

namely, *dl*-alanine, *dl*-α-aminobutyric acid, *dl*-α-aminoenanthic acid, *dl*-α-amino-γ-phenoxybutyric acid, *dl*-C-phenylglycine, *dl*-proline, *dl*-ornithine, *dl*-lysine and 2-aminoindane-2-carboxylic acid, and, in 1949, Gagnon and Nolin (7) synthesized 10 α-amino acids, namely *dl*-α-aminocaprylic acid, *dl*-α-aminopelargonic acid, *dl*-α-aminocapric acid, *dl*-α-aminolauric acid, *dl*-α-(2-ethyl-*n*-butyl)-α-aminoacetic acid, *dl*-α-(2-ethyl-*n*-hexyl)-α-aminoacetic acid, *dl*-α-amino-γ-phenylbutyric acid, *dl*-α-amino-γ-benzylbutyric acid, *dl*-C-cyclohexylglycine, and *dl*-α-amino-β-1-naphthylpropionic acid. To date, 31 α-amino acids have been prepared by the same method. The synthesis of nine other α-amino acids described in this paper illustrates that the Darapsky method for the preparation of these acids is of wide application.

The starting materials, the alkylcyanoacetic esters (I), were obtained by the condensation of organic halides with ethyl cyanoacetate in the presence of sodium ethylate. The yields varied from 20 to 75%.

I II III IV V

R = *n*-nonyl, C_9H_{19}–; *n*-hexadecyl, $C_{16}H_{33}$–; 2-pentyl, C_5H_{11}–; 3-pentyl, C_5H_{11}–; 2-hexyl, C_6H_{13}–; acetyl, CH_3CO–; β-cyclohexylethyl, C_6H_{11}-$(CH_2)_2$–; 1-indanyl, C_9H_9–; γ-chloropropyl, Cl-$(CH_2)_3$–.

Hydrazine hydrate mixed with the esters (I) at room temperature produced the hydrazides (II). These were identified by their condensation products with anisaldehyde.

The azides (III) were formed from the corresponding hydrazides by treatment with nitrous acid, and extracted with ether. The ethyl urethanes (IV) were obtained by refluxing alcoholic solutions of the azides.

The hydrolysis of the urethanes to amino acids (V) was effected by heating with hydrochloric acid. Better yields were obtained in two cases when a mixture of equal volumes of concentrated hydrochloric acid, formic acid (85%), and water was used.

The following α-amino acids were synthesized: *dl*-α-aminoundecylic acid, *dl*-α-aminostearic acid, *dl*-α-amino-β-methylcaproic acid, *dl*-α-amino-β-ethylvaleric acid, *dl*-α-amino-β-methylenanthic acid, *dl*-α-amino-α-acetylacetic acid, *dl*-α-amino-γ-cyclohexylbutyric acid, *dl*-α-(1-indanyl)-α-aminoacetic acid and *dl*-proline.

Experimental*

Ethyl Alkylcyanoacetates (I)

Two equivalents of ethylcyanoacetate were added to a solution of sodium ethoxide (0.5 mole of sodium and 250 ml. of absolute ethyl alcohol). After cooling, one equivalent of organic halide was introduced slowly and the

*All melting points are uncorrected.

TABLE I

ETHYL α-SUBSTITUTED CYANOACETATES, $RCH(CN)COOC_2H_5$

R	Starting material	Yield, %	B.p., ° C.	Formula	Nitrogen, %		$n_D^{°C.}$
					Calc.	Found	
n-Nonyl	$C_9H_{19}Br$	70	153-154 (5 mm.)	$C_{14}H_{25}O_2N$	5.86	5.90	1.4380^{23}
n-Stearyl*	$C_{16}H_{33}Br$	75	167-180 (2 mm.)	$C_{21}H_{39}O_2N$	4.15	4.12	1.446^{27}
2-Pentyl**	$C_5H_{11}Br$	63	101.5 (4 mm.)	$C_{10}H_{17}O_2N$	7.65	7.74	1.4300^{25}
3-Pentyl†	$C_5H_{11}Br$	62	99 (4 mm.)	$C_{10}H_{17}O_2N$	7.65	7.65	1.4335^{20}
2-Hexyl	$C_6H_{13}Br$	50	104-106 (8 mm.)	$C_{11}H_{19}O_2N$	7.11	7.31	1.5310^{25}
β-Cyclohexylethyl	$C_6H_{11}\text{-}(CH_2)_2Br$	70	153-154 (5 mm.)	$C_{13}H_{21}O_2N$	6.28	6.32	1.4588^{24}
1-Indanyl	C_9H_9Br	20	170-175 (8 mm.)	$C_{14}H_{15}O_2N$	6.11	6.50	1.5175^{23}
γ-Chloropropyl	$Br(CH_2)_3Cl$	60	134-136 (6 mm.)	$C_8H_{12}O_2NCl$‡	7.39	7.45	1.4528^{25}

**Ref. (9)* ***Ref. (1)* †*Ref. (2)*
‡*Calc. Cl, 18.7%. Found: Cl, 18.4%.*

reaction mixture boiled under reflux on a water bath until neutral to wet litmus paper. The alcohol was removed by distillation and the residue poured into cold water. The oily layer formed was decanted and the aqueous layer, previously acidified, was extracted several times with ether. The decanted layer and extractions were dried over anhydrous sodium sulphate, the ether evaporated and the residue distilled under reduced pressure.

The properties and yields of the ethyl alkylcyanoacetates are given in Table I.

Alkylcyanoacethydrazides (II)

Equivalent amounts of ethyl alkylcyanoacetates and hydrazine hydrate (100%) were mixed at room temperature. There was evolution of heat and the hydrazides solidified: they were recrystallized from ethanol. α-Cyano-β-methylenanthic hydrazide was obtained as a viscous liquid. The properties of the crystalline products are given in Table II.

TABLE II

α-SUBSTITUTED CYANOACETHYDRAZIDES, $RCH(CN)CONHNH_2$

R	M.p., °C.	Formula	Nitrogen, %	
			Calc.	Found
n-Nonyl	86-87	$C_{12}H_{23}ON_3$	18.7	18.6
n-Stearyl	88-89	$C_{29}H_{57}ON_3$	13.0	12.8
2-Pentyl	140-141	$C_8H_{15}ON_3$	24.9	24.7
3-Pentyl	105-106	$C_8H_{15}ON_3$	24.9	24.7
Acetyl	107-108	$C_5H_9ON_5$	45.2	45.0
β-Cyclohexylethyl	99-100	$C_{11}H_{19}ON_3$	20.1	20.0
1-Indanyl	163-164	$C_{12}H_{13}ON_3$	19.5	19.4
γ-Chloropropyl	102-104	$C_6H_{10}ON_3Cl$*	23.9	23.7

**Calc. Cl, 20.2%. Found: Cl, 19.7%.*

Anisal Derivatives of Alkylcyanoacethydrazides

All the hydrazides were identified by their condensation products with anisaldehyde. The properties of the derivatives are summarized in Table III.

TABLE III

ANISAL DERIVATIVES

Compound	M.p., °C.	Formula	Nitrogen, %	
			Calc.	Found
n-Nonyl	100-101	$C_{20}H_{29}O_2N_3$	12.2	12.2
n-Stearyl	108-109	$C_{27}H_{43}O_2N_3$	9.52	9.70
2-Pentyl	196-197	$C_{16}H_{21}O_2N_3$	14.6	14.9
3-Pentyl	208	$C_{16}H_{21}O_2N_3$	14.6	14.9
2-Hexyl	141-142	$C_{17}H_{23}O_2N_3$	14.0	14.1
Acetyl	208-210	$C_{21}H_{21}O_3N_5$	17.9	18.0
β-Cyclohexylethyl	136-137	$C_{19}H_{25}O_2N_3$	12.8	13.1
1-Indanyl	205	$C_{20}H_{19}O_2N_3$	12.6	12.9
γ-Chloropropyl	136-138	$C_{14}H_{16}O_2N_3Cl$*	14.3	13.9

**Calc. Cl, 12.1%. Found: Cl, 11.8%.*

TABLE IV

AMINO ACIDS AND DERIVATIVES

Compound	Hydrolyzing agent	Time, hr.	Yield, %	Derivative	M.p., ° C.	Formula	Nitrogen, %	
							Calc.	Found
dl-α-Aminoundecylic acid	20% HCl	85	40	Hydantoin	158-159	$C_{12}H_{22}O_2N_2$	12.4	12.3
dl-α-Amino-β-methylcaproic acid	20% HCl	48	60	Copper salt	—	$(C_7H_{14}O_2N)_2Cu$	4.00	4.21
dl-α-Amino-β-ethylvaleric acid	Mixed acids	6	15	Copper salt	—	$(C_7H_{14}O_2N)_2Cu$	4.00	4.07
dl-α-Amino-β-methylenanthic acid	20% HCl	24	30	Copper salt	—	$(C_8H_{16}O_2N)_2Cu$	3.69	3.92
dl-α-Amino-α-acetylacetic acid	20% HCl	48	20	—	—	—	—	—
dl-α-Amino-γ-cyclohexylbutyric acid	20% HCl	48	35	Hydantoin	199-200	$C_{11}H_{18}O_2N_2$	13.3	13.0
dl-α-(1-Indanyl)-α-aminoacetic acid	Mixed acids	4	12	Copper salt	—	$(C_{11}H_{12}O_2N)_2Cu$	3.15	3.40
dl-Proline	20% HCl	48	12	Copper salt	—	$(C_5H_8O_2N)_2Cu$	4.80	4.86

Amino Acids (V)

The original method as outlined by Darapsky (3) was used to transform the substituted cyanoacethydrazides into the azides. Nitrous acid was added dropwise to a cold aqueous suspension of the hydrazides. The azides formed were extracted with ether. Ethanol was added and the ether evaporated. The resulting solutions were refluxed for an hour and the alcohol was removed. The crude urethanes were obtained as pale or dark brown liquids.

Hydrolysis of the urethanes was effected by two different agents : hydrochloric acid (20%) or a mixture of equal volumes of concentrated hydrochloric acid, formic acid, (85%), and water. The amino acids were isolated by the lead oxide method or at their isoelectric point. α-Aminoundecylic acid and α-amino-γ-cyclohexylbutyric acid were precipitated at their isoelectric point.

The properties and yields of the amino acids, together with the properties of their derivatives, are listed in Table IV.

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COMPARISON OF THE NITROLYSIS OF 1, 3-DICYCLOHEXYLIMIDAZOLIDINE WITH THAT OF HEXAMETHYLENETETRAMINE

By J. L. Boivin and George F Wright

Abstract

The nitrolysis of N,N′-dicyclohexylimidazolidine has been effected with formation of the nitrate salt, the nitroso derivative, and the aceto derivative of N,N′-dicyclohexyl-N-nitro-1,2-diaminoethane. The absence of the dinitro derivative suggests that demethylolation is not related to nitramine formation. When formaldehyde splits off to leave a weakly basic amine, as in hexamethylenetetramine nitrolysis, then subsequent nitration will occur; but it will not take place if the resulting amine is strongly basic. Nitrolysis will depend also on reactivity of nitric acid, which can be decreased by addition of ammonium nitrate. This decrease will prevent esterification of intermediate N-methylolamines but not nitrolysis of hexamethylenetetramine types. However both nitrolysis and esterification are thus prevented with dicyclohexylimidazolidine. On the other hand the two nitrolyses are related by the fact that both are accelerated by electropositive chlorine.

The conversion of hexamethylenetetramine to Cyclonite (RDX) by Henning (10) was elaborated by Hale (8) and others (1, 6) so that the reaction was recognized as a scission of a tertiary methyleneamine by nitric acid; that is to say, a nitrolysis (2). One of the ways by which the nitrolysis may be thought to occur is shown as the following ionic mechanism. Hexamine, I, forms a dinitrate, II, stable in aqueous solution. When it is dissolved in acetic anhydride one may consider that a dinitrohexaminium diacetate, III, exists either as an entity or as a path to the final product 1,5-dinitro-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane, DPT. It has not yet been demonstrated whether the formation of such an ionic species or its decomposition is the rate controlling step.

Analogously nitrobenzenediazotic acid, IV, (R = nitrophenyl) may be thought to form a salt with hexamine which would rearrange to di(nitrophenyldiazo)hexaminium hydroxide, V, and thence decompose to 1,5-di-(nitrobenzenediazo)-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane, VI. It is not surprising that this reaction occurs in aqueous solution since nitrobenzenediazotic acid is amphoteric in water and oftentimes reacts as nitrobenzenediazo hydroxide. In contrast, nitric acid will furnish nitronium ion, or a nitronium ion complex such as III, only when the accompanying anion is something other than hydroxyl ion.

Between these two extremes is hypochlorous acid which, functioning slightly as chlorine hydroxide in aqueous solution, may be considered to form with hexamine the ephemeral dichlorohexaminium hydroxide, VIII (via hexaminium dihypochlorite, VII), which would then transform to the actual product, 1,5-dichloro-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane. In all

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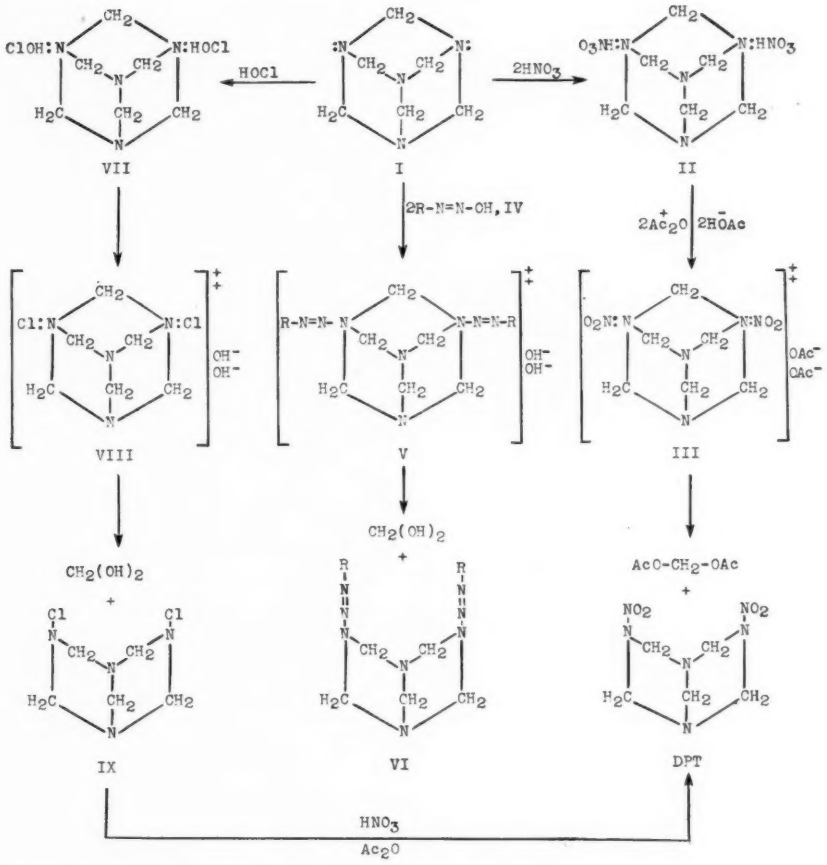


three cases the internal neutralization of the several ions would be effected by scission of methylene di-cation. The latter in water, with two hydroxyl ions, would give formaldehyde hydrate. In acetic acid it would appropriate two acetate anions and thus form the neutral ester, methylene diacetate.

The foregoing explanation has dealt with intermediate complex ions, the existence of which have not been proved rigorously. The suspicion that they exist ephemerally is, of course, based on the known relative stability of nitrobenzenediazonium, electropositive chlorine, and nitronium in decreasing order. Nitrobenzenediazonium is stable in water, while the concentration of nitronium is insignificant except in nonaqueous media. Electropositive chlorine occupies an intermediate position.

The tendency for the electropositive chlorine ion to form an addition product (likewise positively charged) with a neutral molecule in nonaqueous

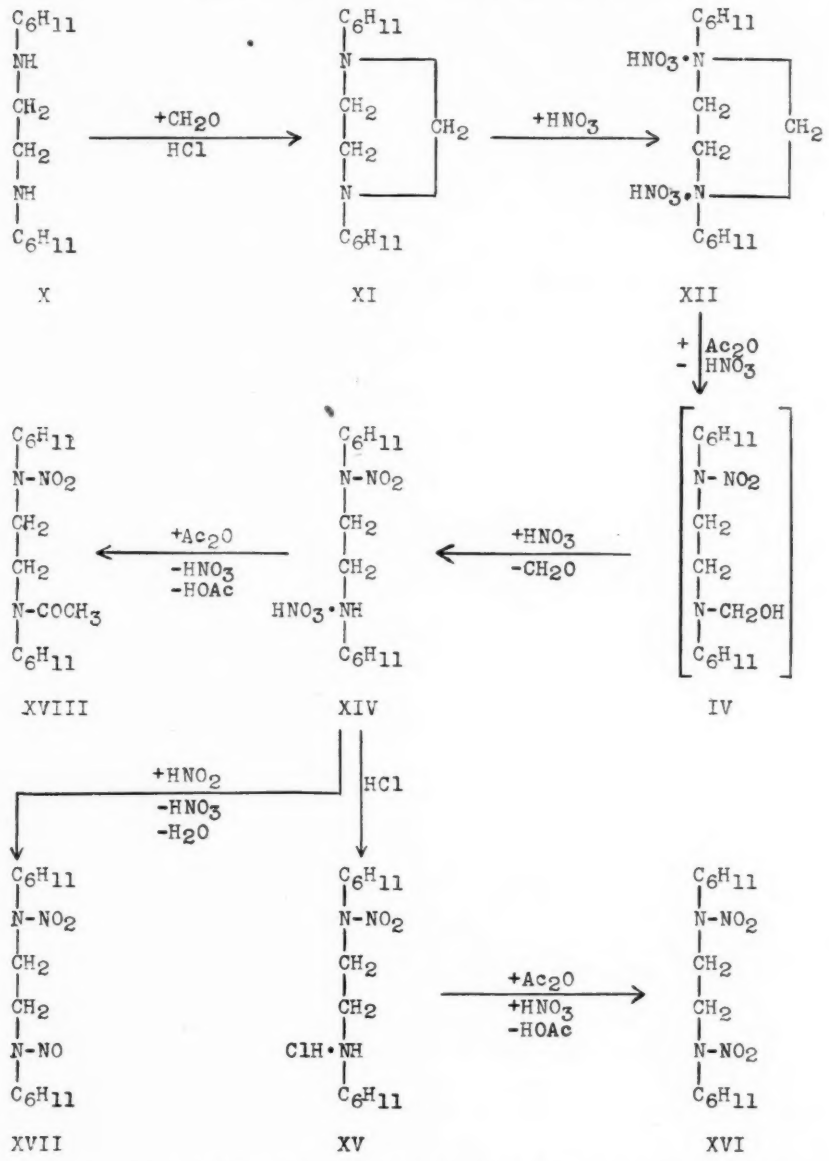
as well as in aqueous media is difficult to discern because of the tendency of the chloramine (which would be formed from it) to act in nonaqueous media as a chlorinating agent, thus to consume itself after it has been formed (14). Despite this instability it has been shown (12) that chloramination of secondary amines occurs with ease in nonaqueous media. In the instance cited, electropositive chlorine acts as a regenerative catalyst for the nitration of secondary amines and the stepwise nature of the reaction has been demonstrated; this implies that the rate of chloramination is more rapid than that of nitramination.

Although this comparison in reaction rate between chloramination and nitramination of secondary amines has thus been demonstrated, there is no published comparable comparison between the relative rates of nitrolysis (scission with nitric acid or its derivatives) and chlorinolysis (scission with hypochlorous acid) of tertiary amines. There was, however a report during the last war (11) that the conversion of hexamethylenetetramine to 1,5-dinitro-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane (2) could be accelerated slightly by addition of hydrogen chloride. In view of a subsequent demonstration that nitric acid will react with hydrogen chloride in acetic anhydride to give an electropositive chlorine derivative (5) it seems possible that when chloride is present the conversion of hexamethylenetetramine to dinitro-*endo*methylenetetrazacycloöctane proceeds via the product of chlorinolysis, IX. Actually we have succeeded in nitrating the supposed intermediate, 1,5-dichloro-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane so that a 48% yield of the dinitro analogue, DPT, is produced.

Neither the nitrolysis nor chlorinolysis of tertiary amines other than hexamethylenetetramine and certain of its derivatives (2,4) has previously been reported. In the present instance we have synthesized 1,3-dicyclohexylimidazolidine, XI, from N,N′-dicyclohexyl-1,2-diaminoethane, X (9). We find that it can be nitrolyzed, although not under conditions favorable for nitrolysis of hexamethylenetetramine.

Thus when XI was treated at 0–25° C with an excess of absolute nitric acid, its dinitrate salt, XII, was formed but no nitrolysis occurred. Subsequent dilution of the solution by water caused hydrolytic scission of this imidazolidine salt, XII, and converted it into the dinitrate salt of N,N′-dicyclohexyl-1,2-diaminoethane. The same salt of X was obtained when the imidazolidine, XI, was treated at 65° C. with a solution of ammonium nitrate in acetic anhydride (with or without nitric acid) and subsequently drowned with water. Under these conditions 40–75% yields of Cyclonite would have been formed from hexamethylenetetramine (6).

When hexamethylenetetramine is treated with nitric acid and acetic anhydride at 25° C. the yield of Cyclonite is insignificant, but nitrolysis does occur. The nitrolysis products are formed in good yield and their structures are known (1, 4). These nitrolyses of hexamethylene are complete in less than one hour, although the reaction mixtures are hetereogeneous.



The treatment of dicyclohexylimidazolidine, XI, with nitric acid and acetic anhydride at 25° C. also precipitated a salt but, in contrast to the rapid reaction with hexamethylenetetramine, the imidazolidine salt redissolved slowly after many hours. Dilution with water caused precipitation of two

new substances. One of these was soluble in boiling water while the other one was not. According to elemental analysis the water-soluble compound was N,N′-dicyclohexyl-N-nitro-1,2-diaminoethane mononitrate, XIV. This designation was proved to be correct by conversion to the corresponding monohydrochloride, XV, which was then nitrated in nitric acid and acetic anhydride to the known N,N′-dicyclohexyl-1,2-dinitraminoethane, XVI.

The second nitrolysis product of N,N′-dicyclohexylimidazolidine mononitrate, was not appreciably soluble in boiling water. It seemed, according to elemental analysis, to be N,N′-dicyclohexyl-N-nitro-N′-nitroso-1,2-diaminoethane, XVII. This structure was confirmed when nitrosation of XIV also yielded the nitro-nitroso derivative, XVII. The yield of XVII from dicyclohexylimidazolidine is variable, depending on reaction time and amount of nitrogen oxides, so that evidently it is formed by nitrosation as a side reaction. This side reaction probably would consume all of the N,N′-dicyclohexyl-N-nitro-1,2-diaminoethane mononitrate eventually if nitrogen oxides were available, despite the fact that nitrosation in acetic anhydride is much slower than in water.

It is evident that N,N′-dicyclohexylimidazolidine, XI, can undergo nitrolysis, initially, like hexamethylenetetramine, though the reaction is much slower. It can be accelerated markedly by increasing the reaction temperature to 60–70° C. However, acetylation also takes place under these conditions since the isolable product is N,N′-dicyclohexyl-N-nitro-N′-aceto-1,2-diaminoethane, XVIII. The identity of this compound has been demonstrated by comparison with the acetylation product of N,N′-dicyclohexyl-N-nitro-1,2-diaminoethane mononitrate.

These examples show that nitrolysis of dicyclohexylimidazolidine differs from that of hexamethylenetetramine in one significant feature. When the latter tertiary amine is nitrolyzed a choice occurs after the initial scission of a methylene-imino linkage. Either the methylol group splits off (with introduction of a second nitro group) or else it is esterified by nitric acid or acetic anhydride. The relative concentration of nitrate or acetate in the medium will determine which of the esters will be formed. On the other hand, esterification has not been observed during the nitrolysis of dicyclohexylimidazolidine under conditions favorable for this reaction. Indeed none of the methylol linkage survives the initial reaction in anhydrous medium. Vacuum distillation of this mixture before it is hydrolyzed transfers all of the formaldehyde into the distillate. No significant amount can be found in the residue after the latter is hydrolyzed with water. Furthermore this demethylolation is not a nitrolysis process since otherwise N,N′-dicyclohexyl-N′,N′dinitraminoethane, XVI, would have been formed.

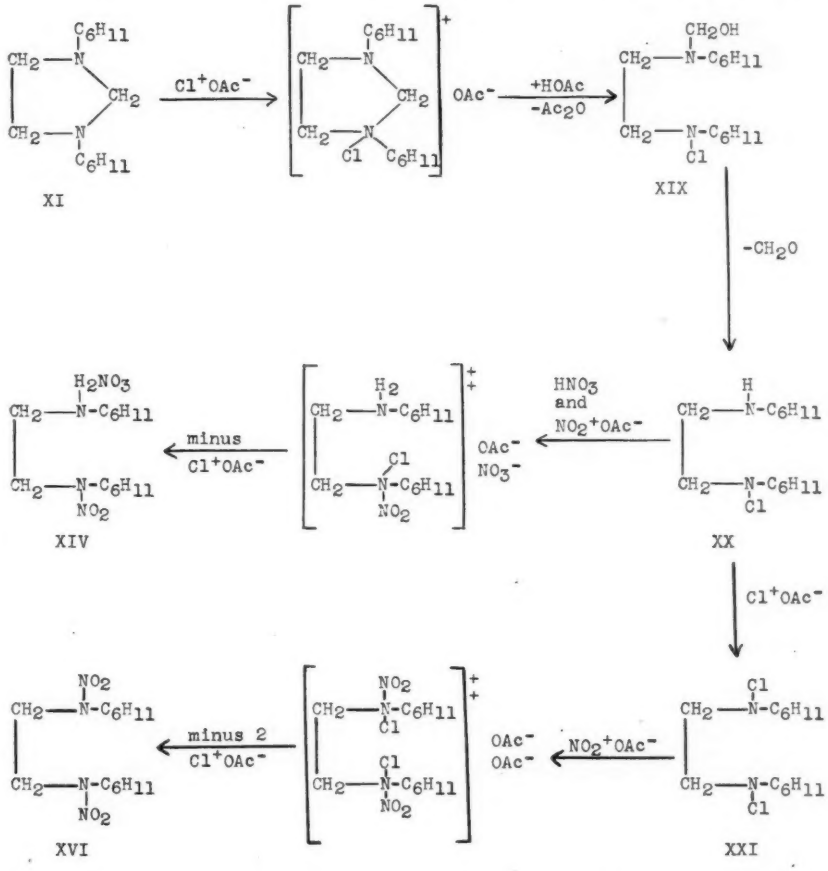
Actually there is no evidence that demethylolation ever involves nitrolysis. To be sure a nitro group appears where formaldehyde has split off during nitrolysis of hexamethylenetetramine, but the two processes are not necessarily

related. The process of demethylolation, after the initial nitrolysis of hexamethylenetetramine, releases a secondary amine group which is so weakly basic that it nitrates spontaneously in presence of nitric acid. Thus 3,5-dinitro-3,5-diazapiperidinium nitrate, which arises as a demethylolation product following nitrolysis of hexamethylenetetramine (13), can be converted with ease to Cyclonite in nitric acid, with or without acetic anhydride. On the other hand the N,N′-dicyclohexyl-N-nitro-1,2-diaminoethane mononitrate, XIV, which results from nitrolysis of N,N′-dicyclohexylimidazolidine cannot be nitrated in these media in absence of chloride catalyst, probably because of its relatively strong basicity (3). From this contrast it seems safe to assume that a methylolamine loses formaldehyde spontaneously in a nitric acid medium, but nitration of the resulting secondary amine is an independent process, the success of which will depend on the ease of nitration of that amine.

This brings into question the action of ammonium nitrate, which completely inhibits nitrolysis of dicyclohexylimidazolidine in nitric acid and acetic anhydride. This might seem in contrast to the effect of ammonium nitrate on the nitrolysis of hexamethylenetetramine, where the desired nitramines rather than the undesirable nitrate esters are obtained by inclusion of the ammonium salt. But in the case of hexamine the nitrolysis is normally a rapid reaction. The esterification reaction, which is relatively slow, may be inhibited by addition of ammonium nitrate to the reaction mixture without decreasing the rate of nitrolysis below a useful limit. Inhibition of esterification in turn allows spontaneous demethylolation to occur, after which nitration will furnish the desired polynitramine. In this sense ammonium nitrate acts by decreasing generally the reactivity of nitric acid. The lesser reactivity will in the case of the more inert dicyclohexylimidazolidine prohibit both esterification and nitrolysis.

It has been shown above in the case of hexamethylenetetramine that electropositive chlorine accelerates nitramine formation. If dicyclohexylimidazolidine is an analogous tertiary amine then nitramine formation in nitric acid–acetic anhydride ought also be enhanced by electropositive chlorine. Actually a fivefold acceleration was observed when a catalytic amount of acetyl chloride was included with the acetic anhydride and nitric acid in order to provide a source of electropositive chlorine. The following mechanism (XI→XIX→XX→XIV) for this reaction is suggested to illustrate the similarity between catalyzed nitration of a secondary amine and catalyzed nitrolysis of a tertiary amine.

In order that these two catalyzed processes might be correlated completely, the stepwise nature of the catalyzed methylene-amine nitrolysis should be demonstrable, like that of the catalyzed amine nitration (12), by isolation of the chloramine followed by its subsequent nitration. The possibility that



XIX or XX could be isolated seemed to be remote, but N,N′-dicyclohexyl-1,2-dichloraminoethane, XXI, might be expected to be stable enough for isolation, if not for characterization. Accordingly dicyclohexylimidazolidine, XI, was treated with aqueous sodium hypochlorite in acetic acid solution. The product was washed with water to remove any amines (as the chlorides and acetates) and then was nitrated in acetic anhydride. The 47% yield of N,N′-dicyclohexyl-1,2-dinitraminoethane, XVI, which was obtained must have been formed from the dichloramine. It has been shown previously (9) that the small amounts of chloride salt which might have been retained in the intermediate could not have produced the dinitro compound in appreciable yield. On these bases it seems reasonable to assume that catalyzed nitrolysis, like catalyzed nitration, proceeds stepwise via an intermediate chloramine.

Experimental*

1,3-Dicyclohexylimidazolidine

A solution of N,N′-dicyclohexyl-1,2-diaminoethane (22.4 gm., 0.1 mole), its dihydrochloride (4.8 gm., 0.02 mole), and formaldehyde (37% in water, 52.5 gm. or 0.6 mole) in 500 ml. absolute ethanol was heated to boiling and then evaporated to dryness under reduced pressure. The residue was made alkaline to pH 8–9 with 10% aqueous sodium hydroxide. The etherous solution, dried over sodium sulphate, was evaporated and the residue fractionally distilled. The fraction boiling at 125–126° C. (0.5 mm.) weighed 26.3 gm. (94% of theoretical). Under 0.7 mm. it boiled at 137–138° C., melting point, + 13 to 15°. Calc. for $C_{15}H_{28}N_2$: C, 76.4; H, 11.9; N, 11.9%. Found : C, 75.2; H, 11.7; N, 12.2%.

When one equivalent of formaldehyde was used the yield was decreased to 25%, and when reaction was carried at 25° C. with five equivalents of formaldehyde the yield was 40%. The compound formed a picrate, crystallizable from nitromethane, which melted at 177–178° C.

1,3-Dicyclohexylimidazolidine Dinitrate

A solution of 5 gm. (0.021 mole) of dicyclohexylimidazolidine in 100 ml. of *absolute* ethanol was chilled to − 60° C. and 2 ml. (0.05 mole) of cold absolute nitric acid was added. The white precipitate was filtered, washed with absolute ethanol and dried at once *in vacuo* since it decomposes in presence of moisture to the dinitrate salt of N,N′-dicyclohexyl-1,2-diaminoethane. It melts initially at 143–144° C. but after two hours this melting point becomes 147–148° C. Analysis for nitric acid (Treadwell ferrous sulphate method) gave 35.2% (calc. 35.1%). Calc. for $C_{15}H_{30}N_4O_6$: C, 49.6; H, 8.37; N, 15.4%. Found : C, 49.2; H, 8.43; N, 15.2%.

N-Nitro-N′-acetyl-N,N′-dicyclohexyl-1,2-diaminoethane

A. From Dicyclohexylimidazolidine Mononitrate

A suspension of 1 gm. (0.0028 mole) of the mononitrate in 10 ml. of acetic anhydride was heated to 60–80° C. for 5–10 minutes. The resulting yellow solution was poured into ice, made basic to pH 9 and extracted quickly with ether. Evaporation of the ether left 0.52 gm., m. p. 90–94° C. This residue was crystallized from hot ethanol. The purified product (m. p. 134° C.) weighed 0.34 gm. or 39% of theoretical yield. It gave a negative Franchimont test for nitramine but a positive lanthanum nitrate test for the acetoxy group. Calc. for $C_{16}H_{29}N_3O_3$: C, 61.8; H, 9.39; N, 13.5%. Found C, 61.5; H, 9.6; N, 13.6%.

When one equivalent of ammonium nitrate was included in the reaction mixture described above, the acetyl compound was not formed, and the

**All melting points have been corrected against reliable standards.*

dinitrate salt of N,N′-dicyclohexyl-1,2-diaminoethane was the only product which could be isolated.

B. From N-Nitro-N,N′-dicyclohexyl-1,2-diaminoethane Mononitrate

When 0.2 gm. (66×10^{-5} mole) of this mononitrate was heated with 5 ml. of acetic anhydride as described above, a yield of 0.16 gm. (86%) of nitroacetyldicyclohexyldiaminoethane, m. p. 131.5–132° C. was obtained after crystallization from hot 50% aqueous ethanol. The compound was also obtained when 0.2 gm. of the mononitrate was treated with 10 ml. of acetic anhydride in 5 ml. of cold dilute aqueous sodium hydroxide initially at pH 9.

N-Nitro-N′-nitroso-N,N′-dicyclohexyl-1,2-diaminoethane

A. From N-Nitro-N,N′-dicyclohexyl-1,2-diaminoethane Mononitrate, XIV

A suspension of 0.33 gm. (0.001 mole) of XIV in 10 cc. water containing 0.69 gm. (0.01 mole) of sodium nitrite was treated at 5° C. with 2 cc. chloroform and then with 0.2 cc. acetic acid over two hours. Then 0.1 cc. of 12% hydrochloric acid was added. After two hours the nonaqueous phase was separated and the aqueous phase twice extracted with chloroform. The combined chloroform phases (4 cc.) were evaporated to dryness and redissolved in 2.5 cc. of boiling benzene. The crystalline product weighed 0.21 gm. (71% of theoretical) and melted at 148.5–149.5° C. A second crystallization from 1.8 cc. of benzene raised this melting point to 149.7–149.9° C. Calc. for $C_{14}H_{26}N_4O_3$: C, 56.4; H, 8.80; N, 18.7%. Found : C, 56.2; H, 8.61; N, 18.4%.

B. From Dicyclohexylimidazolidine Dinitrate, XII

A slurry of 8.47 gm. (0.0268 mole) of the dinitrate in 1.29 cc. (0.03 mole) of absolute nitric acid and 19.4 cc. (0.2 mole) of acetic anhydride was stirred at 27° C. for 14 hr. before the solid was almost dissolved. Then the solution was poured into ice and neutralized with alkali aqueous 30% sodium hydroxide to pH 5. The precipitate was filtered off, weight 8.12 gm., m. p. 135–137° C. This was extracted with 200 cc. of boiling water to leave a residue, m. p. 141–142° C., weight 3.26 gm. This impure nitronitrosodicyclohexyldiaminoethane was crystallized from 40 cc. hot 95% ethanol to yield 2.18 gm., m. p. 148–149° C. or 27% of the theoretical yield.

When the aqueous extract was cooled it precipitated 3.5 gm. (40% of theory) of N-nitro-N,N′-dicyclohexyl-1,2-diaminoethane mononitrate, m. p. 176–177° C.

N-Nitro-N,N′-dicyclohexyl-1,2-diaminoethane Mononitrate

A solution prepared by addition of nitric acid (98% 1.1 ml., 0.026 mole) and acetyl chloride (0.10 gm., 0.0013 mole) in that order to acetic anhydride (17.4 gm., 0.17 mole) and acetic acid (100%, 6.7 ml., 0.117 mole) at 0° C. was stirred at that temperature while the nitrate salt of dicyclohexylimid-

azolidine (4.0 gm., 0.011 mole) was added slowly to form a slurry. After five hours at 25° C. the solid had dissolved. The solution was evaporated, first at 25° C. (18 mm.) and finally at 50° C. (0.1 mm.). The distillate contained much free formaldehyde (2,4-dinitrophenylhydrazone, m. p. 166–167° C.) while the residue contained none.

This residue was dissolved in 25 ml. anhydrous chloroform and the solution diluted with an equal volume of ether. The precipitate, filtered off and ether-washed, weighed 3.3 gm. and melted at 164–165° C. This crude nitrodicyclohexyl-1,2-diaminoethane mononitrate (70% yield) was crystallized from water to melt at 180–181° C.; recrystallization raised this to 181.9° C. Calc. for $C_{14}H_{28}N_4O_5$: C, 50.7; H, 8.59; N, 16.7%. Found : C, 50.8; H, 8.56; N, 16.8%.

The chloroform-ether solution and washings were evaporated to dryness. The residue weighed 0.90 gm. (24% of theoretical) m. p. 130–135° C. Crystallization from ethanol raised this melting point to 150–152° C.; the compound was identified as N,N′-dicyclohexyl-1,2-diacetaminoethane. None of the N,N′-dicyclohexyl-1,2-dinitraminoethane was found.

When the excess of acetic anhydride was replaced by an excess of acetic acid no reaction occurred until acetyl chloride was added, and the dinitrosamine rather than the diacetamide was found to be the chloroform-soluble product. To a solution of 5.9 gm. (0.025 mole) of dicyclohexylimidazolidine in 15 ml. of glacial acetic acid was added 7.28 ml. (0.075 mole) of acetic anhydride proportionately with, but slightly ahead of, 3.46 gm. (0.055 mole) of absolute nitric acid at 25° C. The dicyclohexylimidazolidine mononitrate which was precipitated did not redissolve after 20 hr., and a test portion diluted with water gave no water-insoluble precipitate.

When one equivalent (0.025 mole, 1.96 gm.) of acetyl chloride was added to this mixture at 10° C. the reaction temperature rose to 50° C., then subsided to room temperature over a period of 30 min. The solution was then poured into ice and neutralized to pH 6.5. The insoluble material was filtered off, washed with water, boiled with 10 ml. of water and filtered. The water-insoluble portion was filtered off and crystallized several times from 50% ethanol. It then weighed 1.0 gm. (15% of theoretical) and melted at 144–145° C. It was proved by mixed melting point to be N,N′-dicyclohexyl-1,2-dinitrosaminoethane.

The aqueous filtrate from which the dinitrosamine was separated yielded on cooling 3.1 gm. (41% of theoretical yield) of N-nitro-N,N′-dicyclohexyl-1,2-diaminoethane mononitrate, m. p. 181–182° C.

N-Nitro-N,N′-dicyclohexyl-1,2-diaminoethane Monohydrochloride

A saturated aqueous solution (40 ml.) of the mononitrate was made just basic with cold dilute aqueous alkali and then extracted rapidly with four

5 ml. portions of nitromethane. The combined extracts were evaporated to dryness *in vacuo* and the residue was dissolved in 3 ml. of absolute ethanol. Addition of concentrated hydrochloric acid precipitated 0.15 gm. of the monohydrochloride, m. p. 238–239° C. Two crystallizations from 0.1% hydrochloric acid raised this melting point to 241–242° C. Calc. for $C_{14}H_{28}N_3O_2Cl$: C, 55.1; H, 9.26; N, 13.7%. Found: C, 55.0; H, 9.17; N, 13.4%.

This salt was characterized by treatment of 0.09 gm. (0.003 mole) with 0.246 ml. (0.006 mole) of 95% nitric acid, 0.68 ml. (0.007 mole) of acetic anhydride, and 1 ml. of acetic acid over 24 hr. A 64% yield of N,N′-dinitro-N,N′-dicyclohexyl-1,2-diaminoethane was produced in this manner, and identified by mixed melting point.

Nitration of 1,5-Dichloro-3,7-endomethylene-1,3,5,7-tetrazacycloöctane

To a stirred solution of absolute nitric acid (0.04 mole, 2.52 gm.) and 95% acetic anhydride (0.26 mole, 25 ml.) at 0° C. was added over 10 min. 0.02 mole (3.94 gm.) of dichloro*endo*methylenetetrazacycloöctane (m. p. 79.5° C.). After 30 min. at 0° C. the reaction mixture, now yellow, was warmed to 25° C. over 10 min. and maintained for 30 min. After pouring into ice the solution was neutralized, finally to pH 7. The yield of filtered, water washed 1,5-dinitro-3,7-*endo*methylene-1,3,5,7-tetrazacycloöctane was 2.1 gm., m. p. 200–204° C. after vacuum desiccation at 50° C. Crystallization of this 48% yield raised the melting point to 202–205° C.

Preparation and Nitration of N,N′-Dichloro-N,N′-dicyclohexyl-1,2-diaminoethane

A solution of N,N′-dicyclohexylimidazolidine (5 gm., 0.021 mole) in 100 ml. of glacial acetic acid at 0° C. was treated with 50 ml. of aqueous sodium hypochlorite solution containing 1.5 gm. of electropositive chlorine. After 30 min. the reaction mixture was diluted into ether and the solution washed twice with water, dried with calcium chloride and evaporated *in vacuo*. The semisolid residue was quickly dissolved in 10 ml. of acetic anhydride and added at 0° C. to a chilled mixture of 100% nitric acid (3.15 gm., 0.05 mole) and 30 ml. (0.31 mole) of acetic anhydride. At the beginning of the addition the solution turned yellow; as it warmed to 25° C. a solid began to separate. After 15 hr. at 25° C. this solid was filtered off and washed with ethanol and ethyl ether. The dry weight of this crude N,N′-dicyclohexyl-1,2-dinitraminoethane, m. p. 204–207° C. was 3.1 gm. or 47% of theoretical. After crystallization from nitromethane its melting point (210–211° C.) was not depressed by admixture with an authentic sample (9).

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ACID-BASE REACTIONS IN NONAQUEOUS MEDIA APPLICABLE TO ANALYTICAL AND RADIOCHEMICAL SEPARATIONS

1. REACTIONS IN ORGANIC DERIVATIVES OF AMMONIUM CHLORIDE[1]

By K. Starke

Abstract

Experiments have been performed to investigate whether the common acids formed by water can be replaced by solutions of proton donors in nonaqueous solvents for dissolving metals and oxides. Hydrogen chloride as proton donor and nitrogen compounds as solvents have been examined in particular. It has been found that molten hydrochlorides of some aromatic amines dissolve metals and oxides readily and at convenient temperatures. The advantages to be gained in analytical and radiochemistry by the use of such compounds instead of acids formed by water are discussed.

Introduction

For the preparation and the analysis of inorganic compounds, water is the solvent generally used because of its high dissociating power, its stability, and its cheapness. Its value for dissolving ionized or easily ionizable compounds is unquestioned, and in all cases in which a separation of ions is to be accomplished by formation of insoluble precipitates, water is the most suitable solvent. However, nonionized compounds have become increasingly important in chemistry owing to their volatility and their solubility in organic solvents. These properties frequently make possible a satisfactory separation of two elements if one of them can be introduced into a suitable molecular compound. Nonpolar solvents are used to extract them. However, water is usually present, since it was necessary at the beginning of the procedure for the preparation of the acids needed to transform the metals and oxides into soluble compounds.

The formation of neutral molecules or complex ions is made impossible in many cases in such systems by the high dissociating power of water which may thus become a disadvantage as soon as it is no longer needed for dissolution of the material. Undesired dissociation may even be followed by a reaction with the solvent owing to the fact that water is itself always dissociated. Though hydroxides and oxides formed by hydrolysis are useful in chemical analysis, they are of little value in making good separations of radioactive nuclides. Precipitations will always fail when the solubility product cannot be reached with the extremely small amount of radioactive atoms present. Although another element can be used as a carrier the selectivity of hydroxide precipitations is generally small.

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On the other hand, volatility and solvent extraction methods are successful down to the lowest concentrations. However, as long as water is present, these methods are restricted in use to those molecular compounds which are stable with respect to this polar solvent. It seemed worthwhile to extend these methods with dissolving agents which make possible the use of compounds unstable in the presence of water.

Attempts to prevent decomposition of organic metal complexes in acid solutions by neutralizing them with organic bases led to the discovery that some salts of the latter have a high dissolving power for metals and oxides in the absence of water. This property is to be expected if one considers in detail the phenomenon of formation of acid solutions.

Experimental

No dissolving experiments have been performed with oxygen and sulphur compounds. Reactions of proton donors in water are well known and those reported as taking place in water derivatives and in sulphur compounds indicate little practical value. Reactions of proton donors in liquid ammonia are also known but few attempts have been made to replace it by derivatives of ammonia in order to obtain a nitrogen system which can be used at convenient temperatures Besides an ammonium metal double chloride a number of hydrochlorides of aliphatic and aromatic amines have been tested. Their dissolving powers were checked with metals below the alkali and earth alkali metals in the electromotive force series with emphasis on those giving colored ions as for example uranium, iron, cobalt, copper. In this way reactions taking place, for instance, the formation of complexes, could be easily observed. Only qualitative test tube experiments were carried out in order to survey as many reactions as possible. The report of the results cannot be separated from facts important for this work but already known in other connections. They are, therefore, discussed together in the following sections.

Solutions of Hydrogen Chloride in Oxygen Compounds

Hydrogen chloride was chosen as proton donor because of the high solubility of most chlorides in water and of many of them in organic solvents. The dissolving power of hydrochloric acid is based on the discharge of positive ions formed in the well known equilibrium

$$\begin{array}{ccccccc} H_2O + HCl & \rightleftharpoons & H_2O + H^+ + Cl^- & \rightleftharpoons & H_3O^+ + Cl^- \\ \updownarrow & & & & \updownarrow \\ H_2O.HCl & & & & [H_3O]Cl \end{array} \qquad (1)$$

One of the two possible solid compounds of the components, $[H_3O]Cl$, is unknown, while the other, $H_2O.HCl$, is not stable at room temperature. The fact that hydrogen chloride dissolves in organic oxygen compounds, for example methyl alcohol or ether, suggests analogous equilibria. It has been shown that these solutions attack iron more or less rapidly (3). However,

the formation of water by side reactions can never be excluded with oxygen compounds, and this was the reason they were not investigated. A sulphur system H_2S–HCl analogous to (1) using liquid hydrogen sulphide and hydrogen chloride (8) is not very promising because of the low temperatures which have to be maintained.

Solutions of Hydrogen Chloride in Nitrogen Compounds

In contrast to oxygen containing groups the ammonium ion distinguishes itself by its high stability. In the equilibrium corresponding to (1) stable ammonium chloride is formed:

$$\begin{array}{ccccccc} NH_3 & + & HCl & \rightleftharpoons & NH_3 + H^+ + Cl^- & \rightleftharpoons & NH_4^+ + Cl^- \\ \downarrow\uparrow & & & & & & \downarrow\uparrow \\ NH_3.HCl & & & & & & [NH_4]Cl \end{array} \qquad (2)$$

At atmospheric pressure it sublimes below its boiling point. To obtain a melt which could be used as dissolving agent it would be necessary to maintain an inconveniently high pressure. Nevertheless conditions can be created in which the ammonium ion is discharged like the hydronium ion. It has been shown for instance (4) that solutions of ammonium chloride in liquid ammonia dissolve many metals and oxides and that the reactions are analogous to those in the oxygen system, although in the nitrogen system they frequently result in other end products due to the different solubility of the components in liquid ammonia. Ammonolysis corresponds to hydrolysis and results in formation of amides, imides, and nitrides instead of hydroxides and oxides. However, the experimental inconvenience of maintaining low temperatures prevents a general use of ammonium chloride as solvent for metals and oxides in the liquid system NH_3–HCl.

Reactions of ammonium chloride in the solid state have been reported (5). With the introduction of a third component it can be used at atmospheric pressure and convenient temperatures. By the addition of low melting chlorides a liquid phase can be formed in which the discharge of the NH_4^+ ions takes place as in other fused ammonium salts, for instance, the nitrate (2). For example, ammonium zinc tetrachloride, $(NH_4)_2ZnCl_4$, melts at 150° C. It is used as a soldering salt for removing the oxide layer on metals to be soldered. Experiments showed that the reaction does not always stop after the formation of the chloride, but is followed by ammonolysis ending with precipitation of a nitride.

Comparing the chlorides with regard to their dissolving power one comes to the conclusion that the practical region of usefulness seems to lie between an unstable complex such as $(C_2H_5)_2\,OH^+$ in ether and a stable ion like NH_4^+ which combines with an anion to form a salt. Two series of compounds including either water or ammonia with their organic derivatives can be set up according to the stability of the complex providing the proton to be discharged. While the optimum effect in the series of oxygen compounds lies at the very beginning with water, it is to be expected in the series of

nitrogen compounds somewhere among those derivatives in which the proton is less tightly bound under the influence of the organic radicals replacing the hydrogen atoms of ammonia. From this viewpoint the dissolving power of chlorides of organic nitrogen bases was to be expected and their further investigation seemed to be worth while.

From the great number of organic amines all those had to be eliminated, the hydrochlorides of which have a high melting point, indicating a strong base and promising no better results than ammonium chloride. Furthermore, all compounds which are not easily available in quantity would be of little practical value.

Among the hydrochlorides of aliphatic amines, ethyl ammonium chloride has the relatively low melting point of 109° C. It dissolves metals when heated with them above its melting point, some of them by a vigorous reaction. By addition of xylene, which prevents overheating above its boiling point, the process can be slowed down. Nevertheless, it is harder to get clear solutions with this reagent than with other chlorides mentioned later. As a primary amine the ethylamine being liberated can react with the resulting metal chloride with formation of hydrogen chloride. Furthermore, it forms a strong base and precipitates hydroxides with traces of water. Other hydrochlorides of aliphatic amines gave no better results.

The assumption that chlorides of weak bases should be more suitable has been confirmed with aromatic amines. The hydrochloride of piperidine which forms a rather strong base gives poor results. With pyridinium chloride (melting point 144° C.), quinolinium chloride (135° C.), and anilinium chloride (198° C.), however, the results are more satisfactory, although side reactions cannot always be prevented. While pyridine is more stable against oxidation, aniline is less sensitive to reduction and might sometimes be preferred if metals are to be dissolved.

If no side reactions occur, pyridinium chloride has some advantages over other chlorides. Its low melting point allows the extraction of compounds from its melt by low boiling organic solvents. The pyridine which is liberated during dissolving reactions can be driven off easily if it is not used as solvent for the reaction products. Pyridinium chloride boils at 218°–219° C. and can be purified easily by distillation. Furthermore, the pyridine complex compounds are usually well crystallized and are among the best known in analytical chemistry.

Pyridinium Chloride as Dissolving Agent

The equilibrium between pyridine and hydrogen chloride corresponds to those mentioned above:

$$\begin{array}{ccccccc} C_5H_5N + HCl & \rightleftharpoons & C_5H_5N + H^+ + Cl^- & \rightleftharpoons & C_5H_5NH^+ + Cl^- \\ \downarrow\uparrow & & & & \downarrow\uparrow \\ C_5H_5N.HCl & & & & [C_5H_5NH]Cl \end{array} \qquad (3)$$

Tests were carried out with a number of metals and the results obtained showed that molten pyridinium chloride dissolves many metals more or less quickly according to temperature and electrode potential of the metal. Magnesium, aluminum, and zinc react vigorously with formation of colored reduction products. Uranium, manganese, iron, nickel, tin, lead, and copper dissolve readily, while antimony, bismuth, arsenic, silver, mercury, and platinum show no effect. Thus pyridinium chloride attacks metals above hydrogen in the electromotive force series but, like concentrated hydrochloric acid, copper as well, probably for the reason that the concentration of positive copper ions is small on account of the formation of negative complex ions, making the actual potential of copper far below the normal one. The results are in good agreement with those reported in a paper (1) which came to the author's attention only after this work had been concluded.

The compounds produced by pyridinium chloride depend first on the chemical character of the metal or oxide being dissolved. For instance, alkalis can form only positive ions while elements like platinum are more likely to form negative ions such as $PtCl_6^{2-}$. With other elements lying between these two extremes equilibria exist which can be shifted one way or the other by changing the concentrations of the three components metal chloride–pyridine–hydrogen chloride. For practical reasons the concentrations of the last two are generally equal at the beginning, for this proportion can easily be reached, even in the presence of water, by mixing the two components and driving off the excess of one of them.

Higher concentrations of hydrogen chloride can be produced by the use of HCl gas. A compound C_5H_5N. 2HCl with a melting point of 47° C. is known. Lower concentrations of hydrogen chloride are obtained easily by the addition of pyridine or occur automatically if the dissolution of the metal is performed below the boiling point of pyridine.

The concentration of the metal chloride is limited by its solubility in pyridine. In the system $MeCl_n$–C_5H_5N–HCl the formation of double chlorides or salts of chloroacids is much more frequent than in the corresponding system with water. This prevents the complete transformation of hydrogen chloride into a metal chloride, but it is on the other hand very useful for analytical purposes because of the fact that the crystallizability of complex chlorides is generally superior to that of the simple chlorides.

For the reasons mentioned above, not all proportions $x : y : z$ have to be considered in the equilibrium between the three components:

$$x\,MeCl_n + y\,C_5H_5N + z\,HCl \rightleftharpoons x\,Me^{n+} + y\,C_5H_5NH^+ + (z\text{-}y)H^+ + (nx + z)Cl^- \qquad (4)$$

In the following discussion a tetravalent metal with the co-ordination number 6 is taken as example. The metal used in the tests was uranium.

1. $[MeCl_4] < [C_5H_5N] \approx [HCl]$

If a metal is dissolved in an excess of pyridinium chloride:

$$Me + 4\,C_5H_5NH^+ = Me^{4+} + 4\,C_5H_5N + 2\,H_2 \qquad (5)$$

and if the liberated pyridine is then driven off, the solution will contain mainly pyridinium and chloride ions. As long as the concentration of chloride ions is high compared to that of the metal ions the formation of molecules or even negative complexes is probable:

$$Me^{4+} + 6\,Cl^- \rightleftarrows MeCl_4 + 2Cl^- \rightleftarrows MeCl_6^{2-}. \tag{6}$$

2. $[MeCl_4] \approx [C_5H_5N] \approx [HCl]$

If the dissolution of the metal by pyridinium chloride continues, precipitation occurs either of the metal chloride or of pyridinium metal chloride:

$$MeCl_4 + 2\,C_5H_5NHCl \rightleftarrows (C_5H_5NH)_2MeCl_6. \tag{7}$$

3. $[C_5H_5N] > [MeCl_4] \geqq [HCl]$

After addition of pyridine, the chlorides soluble in it go into solution and dissociation takes place. Metal ions which are formed by shifting the equilibria (7) and (6) to the left side may now be transformed into pyridine complex ions:

$$Me^{4+} + 6\,C_5H_5N \rightleftarrows [Me(C_5H_5N)_6]^{4+}. \tag{8}$$

In contrast to water and ammonia which dissociate causing side reactions:

$$2\,H_2O \rightleftarrows H_3O^+ + OH^- \tag{9}$$
$$2\,NH_3 \rightleftarrows NH_4^+ + NH_2^- \tag{10}$$

such dissociation with pyridine seems to be extremely small. At low temperatures no decomposition of chlorides of weak bases occurs. However, as soon as water is present, precipitation of hydroxides may be caused by the reaction:

$$C_5H_5N + H_2O \rightleftarrows C_5H_5NH^+ + OH^-. \tag{11}$$

4. $[HCl] > [MeCl_4] \approx [C_5H_5N]$

The conditions of the equilibrium with an excess of hydrogen chloride were not investigated. With some metals they may make possible the formation of double chlorides with hydrogen chloride:

$$(C_5H_5NH)_2MeCl_6 + 2\,HCl \rightleftarrows H_2MeCl_6 + 2\,C_5H_5NHCl \tag{12}$$

though it is doubtful if the reaction can be performed quantitatively.

Possible Applications in Analytical and Synthetic Chemistry

Though pyridine is used in analytical chemistry for precipitation of hydroxides, as a solvent, and as a complexing agent it is seldom applied in the complete absence of water, since the latter seemed unavoidable for producing acids. Consequently complexes unstable in the presence of water could not be utilized for analytical purposes. By using pyridinium chloride or analoggous compounds as dissolving agents, however, the equilibrium (6) can generally be shifted farther to the right than in water and its advantage can then be exploited. The pyridinium salts of the chloroacids formed in this way, mainly those of tri- and tetravalent metals, are well crystallized. The precipitates can be made more insoluble by addition of solvents like

chloroform or isoamyl alcohol which are miscible with pyridine and also dissolve the excess of pyridinium chloride leaving behind the complex salt.

The same separation method can be used for mono- and divalent metals which are strongly basic and which either precipitate as simple chlorides or as the basic component of salts of chloroacids. If not already present suitable ions can be added to replace the pyridinium ion in a double salt or chloro-complex. The transformation into ammonium salts was studied because no undesired metal ion has to be introduced by this way. The salts are less soluble in organic solvents such as pyridine and alcohol than the corresponding pyridinium salts and might therefore be useful for analytical separations. They can be formed either by addition of ammonium chloride to the pyridinium chloride before dissolution or by double decomposition:

$$Me + 4\,C_5H_5NHCl + 2NH_4Cl = (NH_4)_2MeCl_6 + 4C_5H_5N + 2\,H_2 \quad (13)$$

$$(C_5H_5NH)_2MeCl_6 + 2\,CH_3COONH_4 = (NH_4)_2MeCl_6 + 2\,C_5H_5N + 2\,CH_3COOH \quad (14)$$

In this connection the reactions with ammonium zinc tetrachloride may be recalled, which allow the formation of zinc double chlorides. Furthermore, it should be possible to transform the chlorides into salts of other strong acids by choosing reagents suitable for these nonaqueous media.

While ionized compounds are precipitated from a mixture of more or less polar solvents containing the excess of pyridinium chloride, molecular compounds on the other hand can be separated by dissolving them in a nonpolar solvent which precipitates pyridinium chloride. Though consisting of neutral nonpolar molecules many chlorides are transformed into ionized chloro-complexes by pyridinium or other chlorides. Therefore, the extraction of simple chlorides by nonpolar solvents is not easily applicable. Nevertheless, it was possible under certain conditions to extract ferric chloride from molten pyridinium chloride by acetone. The solution is poured off after the pyridinium chloride is solidified by cooling and the process can be repeated. In other cases methods have to be found to decompose the chloro-complex.

If an extractable chloride cannot be liberated in this way, the problem arises as to how the metal chlorides, dissolved in pyridinium chloride may be transformed into organic complexes with reagents of the general formula HR. The equilibrium

$$(C_5H_5NH)_2MeCl_6 + 4\,HR + 4C_5H_5N \rightleftharpoons MeR_4 + 6\,C_5H_5NHCl \quad (15)$$

can be shifted to the right by an excess of pyridine which dissociates the chloro-complex provided, of course, that it is not too strong. However, the quantitative separation of the two products from the pyridine solution is difficult. Precipitation of the chloride ions as ammonium chloride under conditions in which the salt is insoluble gives better results:

$$(C_5H_5NH)_2MeCl_6 + 4\,HR + 6\,NH_3 \rightleftharpoons MeR_4 + 2\,C_5H_5N + 6\,NH_4Cl \quad (16)$$

$$(C_5H_5NH)_2MeCl_6 + 4HR + 6\,CH_3COONH_4 \rightleftharpoons MeR_4 + 2\,C_5H_5N + 6\,CH_3COOH + 6\,NH_4Cl \quad (17)$$

Finally the separation of molecular compounds by volatilization has to be considered. Here the chlorides are to be preferred to organic compounds because of their greater stability at higher temperatures. Unfortunately the pyridinium chloro-complexes formed directly cannot be decomposed easily in this way. Though the boiling point of pyridinium chloride is relatively low, it is not possible, at least at atmospheric pressure, to remove this component by distillation without causing side reactions which form colored decomposition products of pyridine. The transformation of pyridinium chloro-complexes into analogous compounds with a more stable basic component by reactions like (13) and (14) seems, therefore, to be necessary. Lithium may be chosen because its chloride has a low melting point and a relatively high boiling point. Thus a second chloride with low boiling point can be separated more easily.

Chlorides made directly by process (7) or by secondary reactions (13) or (14) can serve as nonaqueous electrolytes either in molten state (1) or in pyridine solution for the separation of metals by electrodeposition. It must be taken into account that the electromotive force series in pyridinium chloride is somewhat different from that in water (6).

The stability of negative metal chloro-complexes in molten pyridinium chloride and the fact that they dissociate with different rates on addition of pyridine or water into positive metal ions suggest the use of such solutions in ion exchange columns.

The transition from the water system to the pyridine system is often possible. Water which has to be used for any reason can be removed by distillation if pyridinium chloride is present in high concentration. For the same reason, salts of volatile acids, hydroxides, or oxides formed in the aqueous system are dissolved:

$$Me(OH)_4 + 6\,C_5H_5NHCl = (C_5H_5NH)_2MeCl_6 + 4\,C_5H_5N + 4\,H_2O. \qquad (18)$$

It is worthwhile mentioning that pyridinium chloride can be recovered if the pyridine liberated during dissolution of a metal or its oxide can be used for the precipitation of its hydroxide:

$$Me + 6\,C_5H_5NHCl = (C_5H_5NH)_2MeCl_6 + 4\,C_5H_5N + 2\,H_2 \qquad (19)$$

$$(C_5H_5NH)_2MeCl_6 + 4\,C_5H_5N + 4\,H_2O = 6\,C_5H_5NHCl + Me(OH)_4. \qquad (20)$$

$$Me + 4\,H_2O = Me(OH)_4 + 2\,H_2 \qquad (21)$$

After separation of the hydroxide the excess water is evaporated and the remaining pyridinium chloride can be used over again to dissolve more metal. This process may be useful for enriching trace amounts of an element from large quantities of a metal or an oxide into a given quantity of pyridinium chloride.

The greatest obstacles in the way of a more extended use of pyridine in analytical chemistry lie in the fact that it is relatively unstable, more expensive than water, and toxic. Furthermore, analytical chemistry is built up on the use of water as a solvent for substances of unknown composition. How-

ever, developments in this field have shown that organic solvents and reagents become more important the smaller the quantities of material which have to be dissolved or separated. Here the advantages of organic compounds often outweigh their disadvantages.

Nonaqueous Solutions in Radiochemistry

What is true for micro- and ultramicrochemical quantities of material is also true for the unweighable quantities of radiochemistry. Here, where the conditions are often quite different from those in the chemistry of macro-amounts of material (7) new methods have usually to be developed. With a few exceptions such as fission, the radiochemist has a limited number of elements to separate and their identity is usually known. The quantities to be dissolved and separated are often small, and the compounds to be irradiated can frequently be chosen to suit the separation to follow. The methods described here promise to be useful, therefore, in the preparation of targets to permit a quick separation of the radioactive isotopes. Further acid-base reactions in nonaqueous media which are applicable to analytical and radiochemical separations will be reported in a following paper.

Acknowledgments

I should like to express my appreciation to the National Research Council of Canada for the support of this work.

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SYNTHESIS IN THE SERIES OF *ERYTHRINA* ALKALOIDS[1]

By K. Wiesner, Frank H. Clarke, and S. Kairys

Abstract

One of the four racemates of XIV has been synthesized. The starting material was 2-cyclohexanone acetic acid. This compound was converted into the ethyl ester of 2-aminocyclohexylacetic acid which gave, with homoveratrylbromide, N-homoveratrylhexahydroöxindol. This compound by cyclization and hydrogenation was converted into XIV.

K. Folkers and collaborators have isolated from different species of *Erythrina* two sets of bases: the free alkaloids and the alkaloids liberated on acid hydrolysis (3-10). The investigations of the free bases led to the proposal (9) of a formula which had the common carbon–nitrogen skeleton XV and differed in the number of double bonds and aliphatic methoxyl groups for different alkaloids.

The liberated bases were investigated to a lesser extent, and no structural formula was proposed for them. As it was thought possible that both free and liberated alkaloids have the same carbon–nitrogen skeleton, V, Prelog and his collaborators (15) prepared by hydrogenation, treatment with hydriodic acid and red phosphorus, and methylation with diazomethane from the two groups of alkaloids, two isomeric bases, $C_{18}H_{25}O_2N$.

The two bases were, however, found to be not identical. If the Folkers formula is correct, then the structure of the base derived from the free alkaloid series is represented by XIV.

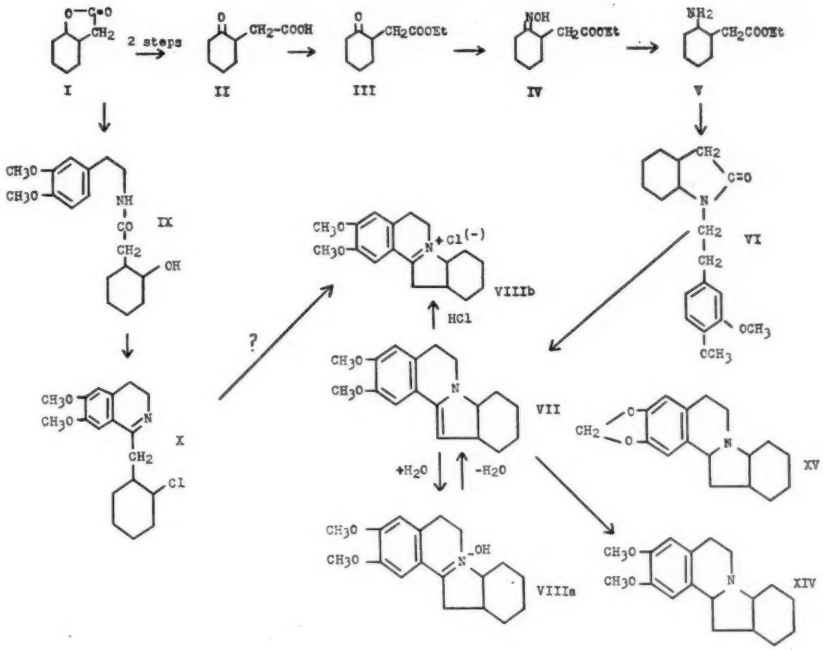
In this communication we describe the synthesis of one of the four possible diastereoisomers of XIV.

The assignment of the spatial configuration, as well as attempts to synthesize the remaining isomers, will form the subject of a subsequent communication.

The starting material was 2-cyclohexanonacetic acid described by Newman and Van der Werf (14). This compound was esterified with ethyl alcohol and hydrochloric acid to III, and the crystalline oxime ester, IV, was prepared.

The oxime ester was hydrogenated in glacial acetic acid with platinum oxide, and it took up in 10–20 hr. the theoretical amount of hydrogen to give the amino ester V. The amino ester was worked up only at room temperature to avoid cyclization. It was however found that it can be quickly distilled in high vacuum, and the analysis, especially the ethoxyl determination, shows that only a very small part has been cyclized in this operation.

[1] *Manuscript received December 30, 1949.*
Contribution from the Department of Chemistry, The University of New Brunswick, Fredericton, N.B.



FORMULA SCHEME I

The complete cyclization to the hexahydroöxindol requires two days or more of refluxing in alcoholic solution. The slowness of this reaction might be an indication of the *trans*-configuration of the amino ester and the resulting cyclization product. We intend to determine the steric configuration by lithium aluminum hydride reduction of the hexahydroöxindol. In the case of a *trans*-configuration the still unknown *trans*-octahydroindol should be formed.

The crude amino ester was refluxed with sodium carbonate and homoveratryl bromide in alcohol for several days, and from the mixture of products N-homoveratrylhexahydroöxindol, VI, was isolated in 44% yield by fractional distillation.

The cyclization of this compound with phosphorus oxychloride gave a crystalline base represented by VII. This base can be recrystallized from ether under nitrogen to a constant melting point 113–115°C. and sublimed in a vacuum, and crystalline derivatives can be prepared from it. If however the recrystallization is performed in hot methanol the solution is colored deeply green and a low yield of a new colorless crystalline compound, m.p. 176°C., is obtained. The analytical results for this compound agree with the formula $C_{18}H_{21}O_4N$. It dissolves very slowly in hydrochloric acid solution and seems to be an oxidation product. It will be investigated later.

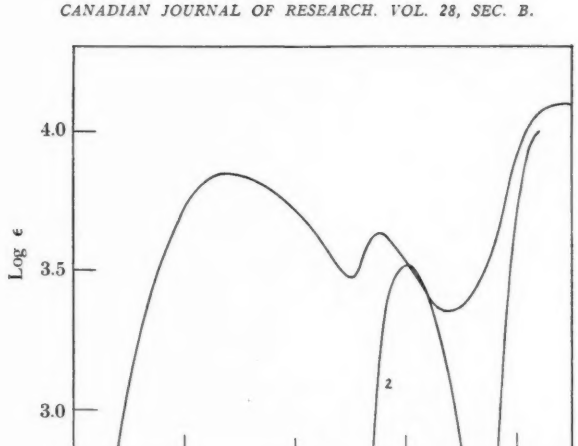


FIG. 1. *Absorption spectra. (1) Compound of melting point 176° C. in alcohol. (2) Compound VI in alcohol.*

The absorption spectra of VII taken in alcohol, alcoholic hydrochloric acid, and sodium hydroxide are shown in Fig. 2. The spectra in acid and neutral medium are identical owing to the fact that the compound in aqueous alcohol is present in the ammonium hydroxide form VIII*a*, which is similar to the hydrochloride VIII*b*. In strongly alkaline medium the compound is present in the form VII, and the spectrum is changed. In contrast to this, the spectrum of the dihydroisoquinoline, X, (Fig. 3) is markedly different in neutral and acid medium (1). This behavior of VII, together with the fact that the

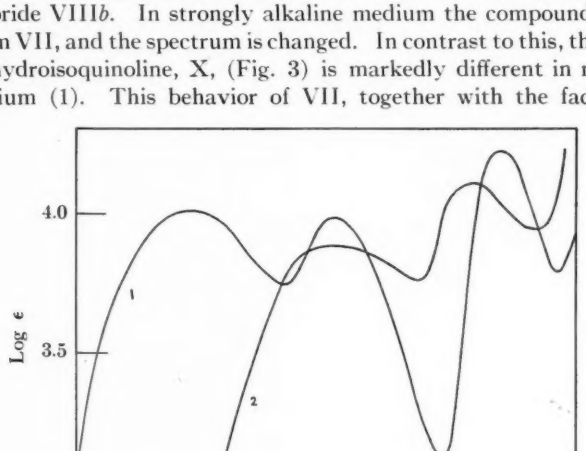


FIG. 2. *Absorption spectrum of Compound VII. (1) In alcohol or alcoholic 0.01 N hydrochloric acid. (2) In alcoholic 0.01 N sodium hydroxide.*

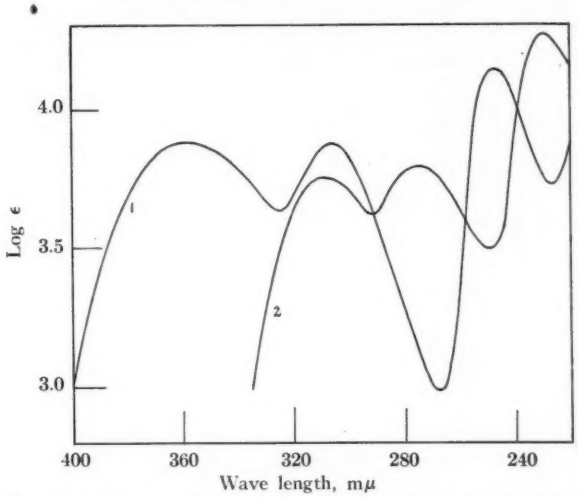


FIG. 3. *Absorption spectrum of Compound X. (1) In alcoholic 0.01 N hydrochloric acid. (2) In alcohol.*

spectrum of this compound in acid and neutral solution is identical with the spectrum of 1 methyl-3, 4-dihydro-6, 7-methylenedioxyisoquinoline (1) in acid solution, is strong support of the structure assigned to it.

The absorption spectrum of the oxidation product of melting point 176°C. is not changed by acid and alkali; it is given in Fig. 1.

VII was hydrogenated in glacial acetic acid with platinum oxide catalyst. It took up 1 mole of hydrogen very rapidly. The reduction product XIV was an oily base which gave a sharply melting homogenous picrate, m.p. 158.5–159.5°C. The absorption spectrum of XIV which is shown in Fig. 4 is identical with the spectrum of the natural base (15).

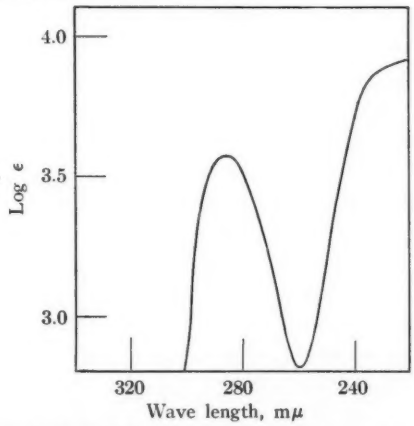


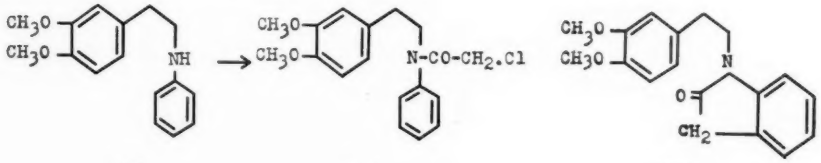
FIG. 4. *Absorption spectrum of the synthetic base XIV.*

In the hope of obtaining the second diastereoisomer of VII, we investigated another synthetic approach to this compound.

Trans-2-oxycyclohexylacetic acid lactone, I, (14) was heated with homoveratrylamine to give the hydroxyamide, IX. This compound was cyclized with phosphorus pentachloride in absolute chloroform in the cold, and at the same time the hydroxy group was replaced by chlorine to give the dihydroisoquinoline, X, characterized as a picrate. It was hoped that by refluxing the base X in alcohol an intramolecular quaternary salt formation would take place leading to VIII*b*.

After two days of refluxing, the base was really transformed into a mixture of hydrochlorides from which a mixture of bases was isolated. The separation of these bases by fractional crystallization of the picrolonates did not succeed. The absorption spectrum of the mixture of bases however showed a band at 360 mμ in neutral alcoholic solution in addition to the normal dihydroisoquinoline spectrum, which disappeared in strongly alkaline solution. This band is probably due to the presence of VII in the mixture of bases, and from the height of the band it is judged that the mixture might contain approximately 10% of this compound.

It is hoped that after reduction it will be possible to separate the tertiary base XIV from the secondary bases.



XI XII XIII

FORMULA SCHEME II

We have further attempted to cyclize N-homoveratryloxindol, XIII, using the same method already described for the hexahydro compound VI. We have attempted the cyclization by using phosphorus oxychloride with and without toluene, phosphorus pentoxide in xylene, and working in an atmosphere of dry nitrogen, but in all cases only amorphous products were obtained. In the synthesis of N-homoveratryloxindol we treated homoveratrylbromide with an excess of aniline in the steam bath and obtained homoveratrylaniline, XI, in good yield. This with chloracetyl chloride gave chloracetylhomoveratrylaniline, XII, which was converted with aluminum chloride to homoveratryloxindole, XIII.

In the experimental part we record the isolation of 2, 2′-dichlordicyclohexyl ether as a byproduct of the oxychlorination of cyclohexene according to the procedure given in Organic Synthesis (12).

This compound has been described by Goldschmidt and Schussler (13) as a product of interaction of chlorine monoxide and cyclohexene in dry chloroform.

Experimental Part*

2-Cyclohexanonacetic acid, II

This was prepared according to the method of Newman and Van der Werf (14) by oxidation of the *trans*-2-hydroxycyclohexylacetic acid with Chromic acid. This in turn has been obtained according to the procedure of the same authors by the reaction of cyclohexene oxide (2, p. 151) with diethylmalonate and saponification of the resulting *trans*-hydroxycyclohexylacetic acid lactone with sodium hydroxide.

The keto acid was characterized by the oxime as well as by the semicarbazone described by Newman and Van der Werf.

The keto acid (0.5 gm.), dissolved in 5 cc. of alcohol, was added to a solution of 0.77 gm. of sodium carbonate and 1 gm. of hydroxylamine hydrochloride in 10 cc. of water and refluxed for two hours. The solution was then evaporated to dryness and extracted with ether. After evaporation of the ether the oily oxime crystallized after long standing. It was recrystallized from ethyl acetate–hexane mixtures. After four recrystallizations it showed a melting point 141–142°C., and it was dried for 24 hr. in a high vacuum at 80°C. for analysis. Calc. for $C_8H_{13}O_3N$: C, 56.12; H, 7.65; N, 8.18%. Found: C, 56.08; H, 7.54; N, 8.42%.

2-Cyclohexanonacetic Acid Ethyl Ester, III

The keto acid (10 gm.) was dissolved in 100 cc. of dry alcohol, and the solution was cooled to 0°C. and saturated with a current of dry hydrogen chloride. After standing overnight the solution was evaporated and worked up by dissolving in ether and washing with water, 5% sodium carbonate, and again with water. The ether was distilled off and the oily ester distilled in vacuum. It boiled at 102–105°C. (3 mm.); yield, 11.1 gm. (94%). Calc. for $C_{10}H_{16}O_3$: C, 65.19; H, 8.75%. Found: C, 65.05; H, 8.67%.

2-Cyclohexanonacetic Acid Ethyl Ester Oxime, IV

The ester (11.1 gm.) was dissolved in about 200 ml. of alcohol. Hydroxylamine hydrochloride (11.1 gm.) and sodium acetate (22.0 gm.) were dissolved in the smallest possible amount of water and added to the solution of the ester.

After refluxing for one hour the alcohol was distilled off under reduced pressure, water was added, and the solution was extracted with ether. The ether layer was washed with 5% sodium carbonate and water, dried and evaporated. The oily oxime ester crystallized from petroleum ether and gave 7.7 gm. (64%) of recrystallized product. The rest was oil which did not crystallize. The compound after three recrystallizations from petroleum ether melted at 60–61°C. and was sublimed for analysis in a high vacuum at 55–60°C. Calc. for $C_{10}H_{17}O_3N$: C, 60.28; H, 8.60%. Found: C, 60.04; H, 8.49%.

**All melting points were determined on a block and are uncorrected.*

2-Aminocyclohexylacetic Acid Ethyl Ester, V

The oxime ester (9.5 gm.) was dissolved in 50 cc. glacial acetic acid and hydrogenated at atmospheric pressure and room temperature with 900 mgm. of prehydrogenated platinum oxide. After 15 to 20 hr., 99% of the theoretical volume of hydrogen had been consumed and the uptake had stopped. The catalyst was filtered and the acetic acid was distilled off in good vacuum at room temperature. The residual oil was worked up with sodium carbonate solution and ether, the aqueous layer being repeatedly extracted with ether as the product seemed to be to some extent soluble in water. The ether extracts were combined, dried, and the ether distilled off in vacuum at room temperature. The oily amino ester weighed 7.4 gm. (84% of the theoretical). A small part was distilled as quickly as possible for analysis. The rest was immediately used for further work.

The analysis, specially the ethoxyl determination, shows that there must have been cyclization to a slight extent but the bulk of the product is the amino ester.

The substance distilled at 100–110°C. (outside temperature) in a collar flask at 1 mm. pressure. Calc. for $C_{10}H_{19}O_2N$: C, 64.70; H, 10.35; OC_2H_5, 24.3%. Found: C, 65.58; H, 10.15; OC_2H_5, 23.37%.

Homoveratrylbromide

This was obtained according to the method of Sugasawa (16) from dimethoxyphenylethyl alcohol and phosphorus tribromide.

Dimethoxyphenylethyl alcohol was obtained from methylhomoveratrate according to the procedure of Fulton and Robinson (11). By the careful use of dried alcohol (diethylphthalate method) and careful exclusion of moisture during the reaction the yield of the Bouveault–Blanc reduction was raised to 70% even in small runs.

N-Homoveratrylhexahydroöxindol, VI

The crude amino ester (7.4 gm.) and homoveratryl bromide (4.9 gm.) were dissolved in 200 ml. of absolute alcohol. To the solution was added 5 gm. of anhydrous sodium carbonate and it was refluxed for 64 hr. Small portions of sodium carbonate were added at intervals of 12 hr. The alcohol was then distilled off, water was added, and the product was extracted with ether. The ether was dried and evaporated, leaving 8.2 gm. of a brownish oil. This was distilled in a Hickman flask at 0.1 mm. The first fraction, boiling at 88°C. to 92°C., weighed 4.6 gm.

The second fraction, containing the desired product, boiled from 172° to 177°C. (mostly at 175°C.); yield, 2.7 gm. (44.5%). It was dissolved in ether and shaken with dilute hydrochloric acid. In this operation a very small amount (*ca.* 7%) of bases was removed. Part of the product was distilled in a

collar flask for analysis. It is a colorless oil which becomes waxy after cooling. Calc. for $C_{18}H_{25}O_3N$: C, 71.25; H, 8.31; N, 4.62%. Found: C, 70.94; H, 8.28; N, 4.44%. The absorption spectrum is shown in Fig. 1.

Cyclization to VII

The above compound (0.625 gm.) was dissolved in 4 ml. absolute toluene and 1 ml. of phosphorus oxychloride. Dry nitrogen was passed through the flask and the mixture was refluxed for 1½ hours.

After cooling, 30 ml. of petroleum ether was added and the solvent poured off the brownish-red oil. The oil was dissolved in 0.2 *M* hydrochloric acid and extracted three times with ether. It was made alkaline with 10% sodium hydroxide and the precipitated base extracted with ether. The ether solution was dried and evaporated in a vacuum. In the flask remained 0.361 gm. of slightly yellowish crystals, m.p. 107–112°C.

They were recrystallized from ether under nitrogen. After four recrystallizations they showed the constant melting point 113–115°C.

For analysis one sample was dried in a high vacuum at 50°C. for 24 hr.; another was sublimed in a high vacuum at 100°C. Calc. for $C_{18}H_{23}O_2N$: C, 75.76; H, 8.12; N, 4.91%. Found: C, 76.03; H, 8.25; N, 5.10, 4.98%.

The hydrochloride of the base can be precipitated from ether solution with dry hydrogen chloride and forms a hygroscopic oil which crystallizes after long standing in the ice box.

The *picrolonate* was prepared by adding the base dissolved in ether to a methanolic solution of the calculated amount of picrolonic acid. After four recrystallizations from methanol–ether it melted at 174–176°C. Calc. for $C_{28}H_{31}O_7N_5$: C, 61.20; H, 5.68%. Found: C, 61.31; H, 6.12%.

Hydrogenation to XIV

The recrystallized base VII (329 mgm.) was hydrogenated in glacial acetic acid with prehydrogenated platinum oxide (65 mgm.). The uptake was completed in three hours and was 30.3 ml. (calc. for 1 mole, 28.4 ml.).

The base was isolated in the usual way; it was a colorless oil. Yield, 320 mgm.

The oil was converted into the picrate which formed yellow needles; m.p. 145–155°C. After three recrystallizations the melting point was 158.5–159.5°C.* and it was not changed by further crystallizations.

For analysis the picrate was dried in a high vacuum at 60° for 24 hr. Calc. for $C_{24}H_{28}O_9N_4$: C, 55.81; H, 5.46; N, 10.85%. Found: C, 55.89; H, 5.47; N, 10.80%: C, 55.99; H, 5.54; N, 10.93%.

**Decomposition.*

Oxidation of the Base VII

If the base (m.p. 113–115°C.) is recrystallized in hot methanol without protection from the air the solution is colored deep green, and after concentration of the solution a colorless crystalline compound precipitates in a 30% yield. It is sparingly soluble and can be easily recrystallized from methanol. It melts with decomposition at 173–176°C. and on exposure to sunlight is colored red on the surface. A more convenient way to obtain this substance is to dissolve the base (m.p. 113–115°C.) in chloroform and expose it for a day to the air. The solution is colored deep red after this period and the compound (m.p. 176°C.*) can be separated from the unchanged base (m.p. 115°C.) and the red coloring matter by chromatography on aluminum oxide. The compound (m.p. 176°C.*) is eluted with 1% methanol in chloroform. For analysis it was recrystallized several times from methanol and dried in a high vacuum at 60°C. for 24 hr. Calc. for $C_{18}H_{21}O_4N$: C, 68.55; H, 6.67; OCH_3, 19.65. Found: C, 68.42; H, 6.60: C, 68.11; H, 6.79; OCH_3, 20.48%.

Trans-2-oxycyclohexylacetic Acid Homoveratrylamide, IX

Homoveratrylamine (14.7 gm.) and *trans*-2-oxycyclohexylacetic acid lactone (12.5 gm.) (14) were heated under nitrogen to 200°C. for one hour. After cooling, the mixture crystallized, and it was recrystallized from ethyl acetate to give 16.9 gm. of nearly pure product. After five recrystallizations the substance melted at 109.5–110.5°C.; it was dried at 85° in a high vacuum for 24 hr. for analysis. Calc. for $C_{18}H_{27}O_4N$: C, 67.3; H, 8.5; N, 4.36; OCH_3, 19.3%. Found: C, 67.19; H, 8.33; N, 4.68; OCH_3, 18.76%.

Ring closure to the dihydroisoquinoline, X

The amide (0.729 gm.) and 1.36 gm. of phosphorus pentachloride were dissolved in 100 cc. of absolute chloroform and left at room temperature for 50 hr. After this time the chloroform was evaporated in a vacuum and the residue dissolved in approximately 100 cc. of warm water. The aqueous solution was then cooled with ice and made alkaline with 10% sodium hydroxide. The precipitated base was extracted with ether and the ether solution after drying was evaporated in a vacuum at room temperature. Yield; 0.7 gm. of a yellowish oil which on very slight warming was decomposed and converted into the hydrochloride of a new base.

It was therefore characterized as picrate, which is quite stable. It melted, after four recrystallizations from methanol, at 185–187°C. Calc. for $C_{24}H_{27}O_9N_4Cl$: C, 52.32; H, 4.94; N, 10.17; OCH_3, 11.26; Cl, 6.46%. Found: C, 52.31; H, 4.95; N, 10.23; OCH_3, 11.49; Cl, 7.13%.

For the determination of the absorption spectrum the base was recovered from the picrate.

**Decomposition.*

*Homoveratrylaniline, XI**

Homoveratrylbromide (5 gm.) and 12 gm. of aniline were heated for three hours in a boiling water bath. Excess of aniline was removed by steam distillation, the solution made basic, and the product taken up in ether. The ether was evaporated and the product distilled from a Hickman flask in a high vacuum. Yield, 4.23 gm. (80%).

For analysis the product was redistilled from a collar flask. Boiling point (outside) 175–180°C. (0.1 mm.). Calc. for $C_{16}H_{19}O_2N$: C, 74.71; H, 7.39; N, 5.45%. Found: C, 74.67; H, 7.17; N, 5.71%.

*Chloracetylhomoveratrylaniline, XII**

Homoveratrylaniline (4.65 gm.) was dissolved in dry ether and treated with an ether solution of 2 gm. of chloracetylchloride.

The product separated as an oil. Ether was removed by distillation, and the product dissolved in chloroform and washed with dilute hydrochloric acid and water. After drying, the solvent was evaporated and the residual oil dried in a vacuum overnight. Yield; 95.8% of crude product which crystallized slowly (m.p. 66–68°); this was immediately used for the further reaction.

Homoveratryloxindol, XIII

The crude chloracetyl compound (5.5 gm.) was dissolved in dry carbon bisulphide, 10 gm. of powdered aluminum chloride was suspended in the solution and the solvent evaporated. The mixture was heated for 10 hr. to 160–170° in an oil bath. After this the mixture was decomposed with dilute hydrochloride acid and ice. The black tarry product was purified by repeatedly extracting with hot water, from which it separated after cooling in a comparatively pure crystalline state. Yield, 2.6 gm.

It was crystallized first from chloroform and then from ethanol. It forms beautiful white needles, m.p. 227.5–229°C. It was dried in a high vacuum at 60°C. for 24 hr. Calc. for $C_{18}H_{19}O_3N$: C, 72.7; H, 6.46; N, 4.71%. Found: C, 72.55; H, 6.29; N, 4.48%.

2, 2′-Dichlorodicyclohexyl Ether

The oxychlorination of cyclohexene was performed in several runs exactly according to the procedure in "Organic Syntheses" (12, p. 31). All runs gave the expected yield of approximately 70% 2-chlorocyclohexanol. The distillation residues crystallized accidentally after standing for many weeks at a low temperature. After five recrystallizations from methanol the compound melted at 70–71°C. It was sublimed for analysis in a high vacuum at 50–60°C. Calc. for $C_{12}H_{20}OCl_2$: C, 57.38; H, 8.05; Cl, 28.22%. Found: C, 57.28; H, 8.25; Cl, 28.63. The compound had no active hydrogen. 2, 2′-Dichlorodicyclohexylether is reported by Goldschmidt and Schussler to melt at 69°C.

**This compound was prepared by W. A. Edmiston.*

Acknowledgments

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We wish also to express our gratitude to Dr. F. J. Toole for his generous support and interest in our research work.

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Notes on the Preparation of Copy

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